



This is a digital copy of a book that was preserved for generations on library shelves before it was carefully scanned by Google as part of a project to make the world's books discoverable online.

It has survived long enough for the copyright to expire and the book to enter the public domain. A public domain book is one that was never subject to copyright or whose legal copyright term has expired. Whether a book is in the public domain may vary country to country. Public domain books are our gateways to the past, representing a wealth of history, culture and knowledge that's often difficult to discover.

Marks, notations and other marginalia present in the original volume will appear in this file - a reminder of this book's long journey from the publisher to a library and finally to you.

### Usage guidelines

Google is proud to partner with libraries to digitize public domain materials and make them widely accessible. Public domain books belong to the public and we are merely their custodians. Nevertheless, this work is expensive, so in order to keep providing this resource, we have taken steps to prevent abuse by commercial parties, including placing technical restrictions on automated querying.

We also ask that you:

- + *Make non-commercial use of the files* We designed Google Book Search for use by individuals, and we request that you use these files for personal, non-commercial purposes.
- + *Refrain from automated querying* Do not send automated queries of any sort to Google's system: If you are conducting research on machine translation, optical character recognition or other areas where access to a large amount of text is helpful, please contact us. We encourage the use of public domain materials for these purposes and may be able to help.
- + *Maintain attribution* The Google "watermark" you see on each file is essential for informing people about this project and helping them find additional materials through Google Book Search. Please do not remove it.
- + *Keep it legal* Whatever your use, remember that you are responsible for ensuring that what you are doing is legal. Do not assume that just because we believe a book is in the public domain for users in the United States, that the work is also in the public domain for users in other countries. Whether a book is still in copyright varies from country to country, and we can't offer guidance on whether any specific use of any specific book is allowed. Please do not assume that a book's appearance in Google Book Search means it can be used in any manner anywhere in the world. Copyright infringement liability can be quite severe.

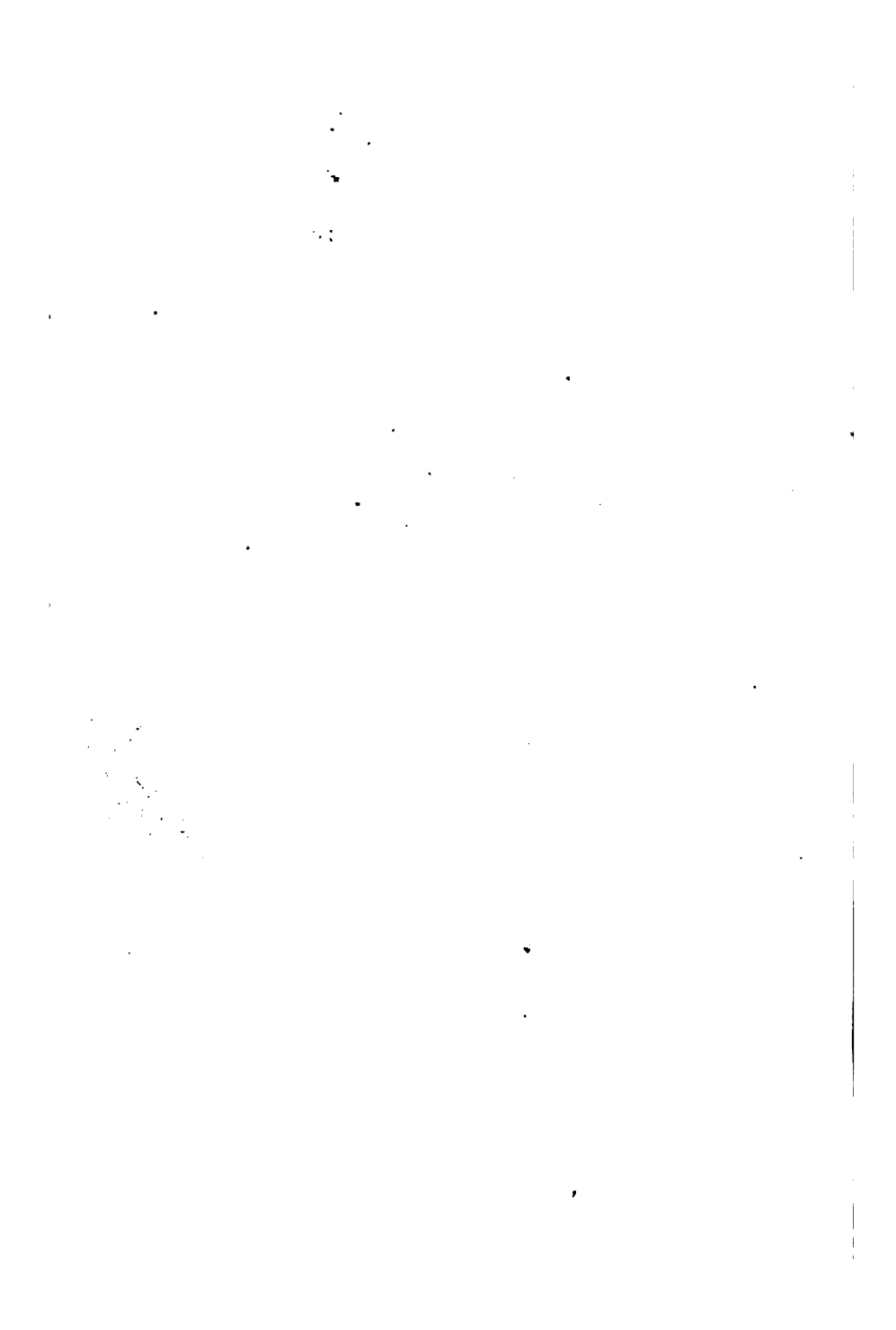
### About Google Book Search

Google's mission is to organize the world's information and to make it universally accessible and useful. Google Book Search helps readers discover the world's books while helping authors and publishers reach new audiences. You can search through the full text of this book on the web at <http://books.google.com/>

CLASS BOOK  
OF  
INORGANIC CHEMISTRY  
BY  
D. MORRIS. B.A.







A CLASS-BOOK  
OF  
INORGANIC CHEMISTRY:  
WITH TABLES OF  
CHEMICAL ANALYSIS,  
AND DIRECTIONS FOR THEIR USE:  
COMPILED SPECIALLY FOR  
PUPILS PREPARING FOR THE OXFORD AND CAMBRIDGE MIDDLE-CLASS  
EXAMINATIONS, AND THE MATRICULATION EXAMINATION OF  
THE UNIVERSITY OF LONDON.

BY  
D. MORRIS, B.A.  
TEACHER OF CHEMISTRY IN LIVERPOOL COLLEGE.



LONDON:  
GEORGE PHILIP AND SON, 32, FLEET STREET.  
LIVERPOOL: CAXTON BUILDINGS, SOUTH JOHN STREET,  
AND 49 AND 51, SOUTH CASTLE STREET.

1870.

193. f. 31.



## P R E F A C E .

---

THE following pages are simply enlarged notes of the most important facts of Inorganic Chemistry, which were originally culled from the best modern books on this subject for the use of the pupils under my charge. I have long felt the need of an Elementary Text-Book which should satisfy the requirements of the "Middle-Class Examinations;" and, to meet this want, I have compiled, in addition to the main facts of the science, Tables of Analysis of simple compound substances, with directions for their use, and Rules for sundry Arithmetical Exercises.

I strongly recommend to beginners in Qualitative Analysis the use of Table on page 148, on account of the simplicity of the method, as well as for the reason there stated.

The numerous Questions found in the work have been selected from the examination papers of Oxford, Cambridge, and London Universities, and the Science and Art Departments. The questions of the latter are denoted by \*: those of the Universities, by the initial letter of each respectively.

My thanks are due to Dr. Carter, of the Liverpool School of Science, for many valuable hints and suggestions.

D. M.

LIVERPOOL,

Aug., 1870.





# CONTENTS.

---

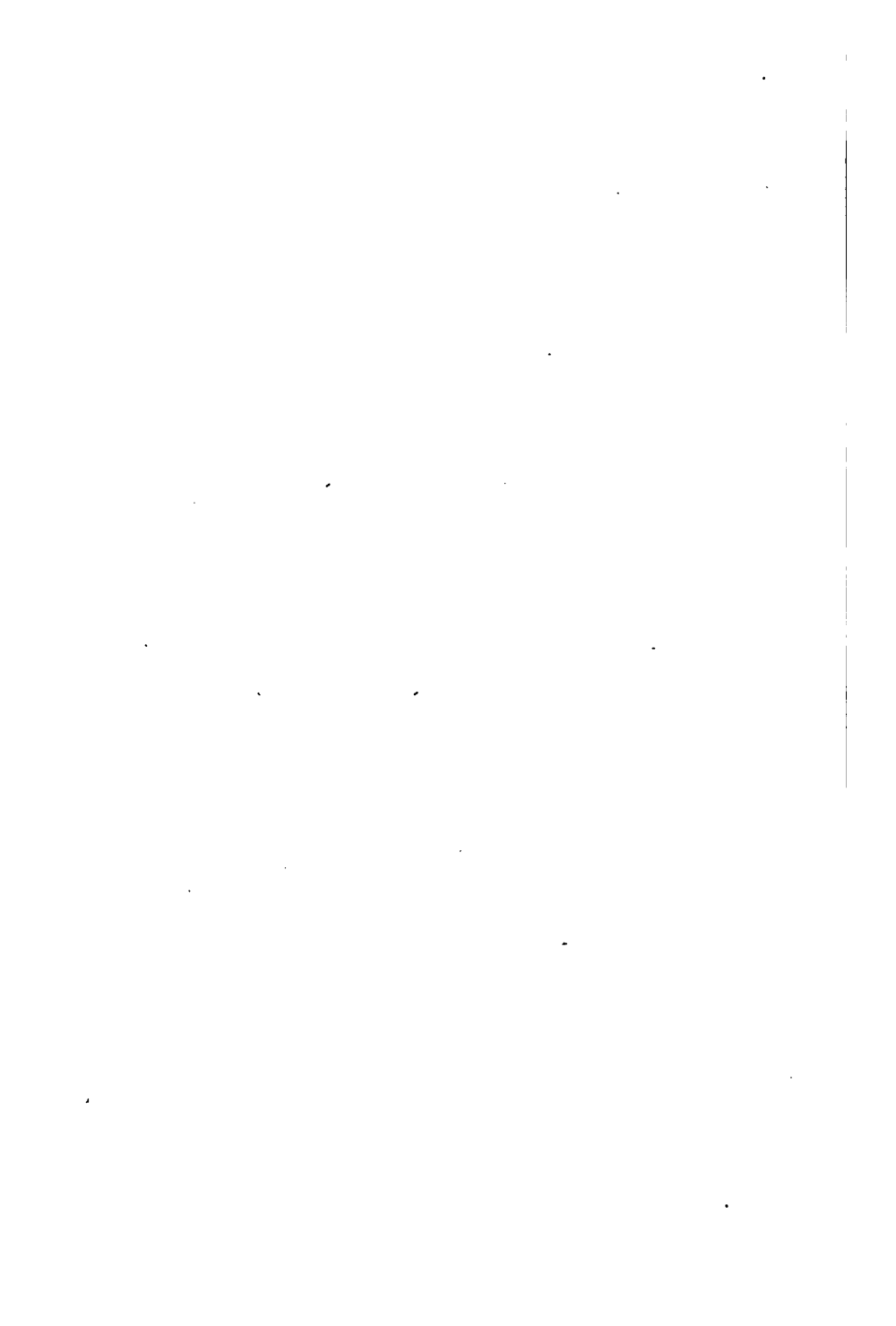
	PAGE.
Introduction . . . . .	1
Oxygen . . . . .	6
Measurement of temperature of bodies . . . . .	10
Expansion of gases . . . . .	11
Questions on Oxygen, &c. . . . .	12, 13
Hydrogen . . . . .	13
Compounds of Oxygen and Hydrogen . . . . .	15
Measurement and weight of gaseous volumes . . . . .	17
Questions on Hydrogen . . . . .	18
Nitrogen . . . . .	19
Atmosphere . . . . .	20
Compounds of Nitrogen and Oxygen . . . . .	22
Questions on Nitrogen, &c. . . . .	26, 27
Carbon . . . . .	27
Compounds of Carbon and Oxygen . . . . .	29
Questions on Carbon, &c. . . . .	32
Compounds of Carbon and Hydrogen . . . . .	33
Nature and structure of Flame . . . . .	35
Davy Lamp . . . . .	36
The Blow-pipe . . . . .	37
Questions on Compounds of Carbon and Hydrogen, &c. . . . .	38
Chlorine . . . . .	38
Compounds of Chlorine and Oxygen . . . . .	41
"                "                Nitrogen . . . . .	43
"                "                Phosphorus . . . . .	43
Questions on Chlorine and its Compounds . . . . .	43, 44
Bromine . . . . .	44

	PAGE.
Iodine . . . . .	45
Fluorine . . . . .	47
Sulphur . . . . .	48
Compounds of Sulphur and Oxygen . . . . .	49
"                    Hydrogen . . . . .	52
Questions on Sulphur, &c. . . . .	54
Selenium . . . . .	54
Silicon . . . . .	56
Boron . . . . .	57
Phosphorus . . . . .	58
Compounds of Phosphorus and Oxygen . . . . .	60
Phosphuretted Hydrogen . . . . .	61
Questions on Phosphorus, &c. . . . .	62, 63
 THE METALLIC ELEMENTS . . . . .	 64
Chemical properties of Metals . . . . .	66
Distribution, &c., of Metals . . . . .	67
METALS OF THE ALKALIS :—	
Potassium . . . . .	68
Sodium . . . . .	71
Lithium, Cæsium, Rubidium . . . . .	74
Ammonium . . . . .	75
METALS OF THE ALKALINE EARTHS :—	
Barium . . . . .	78
Strontium . . . . .	79
Calcium . . . . .	80
Magnesium . . . . .	83
Rules for sundry Arithmetical Exercises . . . . .	85
Questions on Alkalis and Alkaline Earths . . . . .	86
METALS OF THE EARTHS :—	
Aluminum . . . . .	88
METALS PROPER :—	
Zinc . . . . .	91
Iron . . . . .	98
Manganese . . . . .	98
Cobalt . . . . .	101
Nickel . . . . .	102
Chromium . . . . .	103

# CONTENTS.

vii

	PAGE.
Copper . . . . .	106
Lead . . . . .	108
Silver . . . . .	111
Mercury . . . . .	114
Bismuth . . . . .	117
Cadmium . . . . .	118
Gold . . . . .	119
Platinum . . . . .	120
Tin . . . . .	122
Antimony . . . . .	124
Arsenic . . . . .	126
Rare Metals, as Uranium, &c. . . . .	129
Questions on Metals Proper . . . . .	132
Hints on Qualitative Analysis . . . . .	135
List of Chemicals for Analysis . . . . .	137
Directions for Analysing Simple Compounds . . . . .	139
Tables of Analysis . . . . .	140
Tables of Weights and Measures . . . . .	149
Index . . . . .	151



# INTRODUCTION.

---

**CHEMISTRY** is the science which treats of the nature and properties of all things which enter into the composition of earth, sea, air, plants, and animals.

Nearly all the visible objects of the world are **compound** bodies. These compound bodies are capable of being broken up into simpler forms. This process is called **decomposition**. When a body resists all further decomposition, it is called an **element**. Water, for example, consists of two gases, oxygen and hydrogen; but neither of these gases can be resolved into anything simpler, therefore each of them is called an element.

The number of known elements is **sixty-four**. Of these, **fifty-one** are **metals**, and **thirteen** non-metallic substances.

The science of Chemistry is arranged in two great divisions, called **Organic** and **Inorganic**. Organic Chemistry is the chemistry of plants and animals. Inorganic Chemistry treats of the nature of substances which do not originate in animal and vegetable life. This latter division, therefore, includes all the elements.

Before entering upon the discussion of these, it is necessary to have a clear notion of the **weight** of substances. **Weight** is the pressure downwards in obedience to the law of **gravity**; but chemistry considers weight in three different relations, as

**absolute weight, specific weight, and atomic weight.** **Absolute weight** is the amount of weighable matter in a body. This was usually expressed, by English Chemists, in Avoirdupois weight, the grain being taken as a standard; but the **Metric system**, or French Decimal system of weights, is now more commonly used. Tables of English and French weights and measures will be found in the Appendix.

**Specific weight, or specific gravity**, is the absolute weight of a body compared with the absolute weight of another body of the **same bulk**, taken as a standard. **Water**, at a certain temperature (60° Fahr.) (15°·5 C.), is the standard of comparison for liquids and solids. **Air** is the standard for gases in this book. Hydrogen may also be taken as the standard of gases; in that case the sp. gr. of the gas and its atomic weight will be identical. Thus, if we take equal bulks of iron and water, and compare their weights, we shall find that the iron is 7·8 times heavier than the water; therefore, we say the specific gravity (sp. gr.) of iron is 7·8,—water being taken as unity.

The sp. gr. of a **liquid** is found by weighing equal bulks of the liquid and water at the standard temperature, and then dividing the weight of the liquid by the weight of the water,—the quotient will be the sp. gr.

The sp. gr. of a **solid** may be found on the same principles. Archimedes, a philosopher of Syracuse, who lived 287–212 B.C., discovered that when a solid is immersed in a fluid, it loses a portion of its weight, and this loss is equal to the weight of the fluid it displaces; that is, the loss of weight of the solid is the weight of an equal bulk of the fluid. From this we have the following rule for finding the sp. gr. of a solid. “Weigh the solid in air, then weigh it in water; divide the weight in air by the loss of weight in water, and the quotient will be the sp. gr.”

Specific gravity is also called **density**; but density really means the **mass or quantity of matter** in a body.

**Atomic Weight** is usually defined as the proportion by weight with which bodies chemically unite with one another. Each element is supposed to be made up of indivisible particles called **atoms**. An atom may be defined to be the smallest proportional weight of an elementary substance which is capable of existing in **chemical combination**. These atoms unite to form compounds; but each atom is always of the same weight, and always unites in the same proportion by weight. An atom of Hydrogen is taken as the standard of comparison, and its atomic weight is called 1. An atom of Oxygen weighs sixteen times heavier than an atom of Hydrogen. An atom of Nitrogen weighs fourteen times heavier; therefore we call the atomic weights of Oxygen and Nitrogen 16 and 14 respectively. The atomic weight is also called **combining weight**, or **combining proportion**, and **relative weight**. **Chemical equivalent** is also used to express the same thing, but this expression is not quite accurate.

The smallest cluster of atoms capable of existence by itself is called a **Molecule**. Thus, when we speak of a molecule of hydrogen, we mean two atoms of the gas, because this is the smallest amount which can exist in a separate or uncombined state.

Every element is distinguished by one or two letters, called its **symbol**. This symbol is usually the first letter of the name of the element. If two letters are needed, the first is a capital letter, and the second a small one. If two or more elements combine to form a compound body, the symbol of each is written in capital letters. Thus, water, a compound of Hydrogen (H), and Oxygen (O), is written  $H_2O$ . The symbol of a substance, besides, represents the combining proportion of the element. Thus, O stands not only for the element Oxygen, but also for its relative weight, 16, as compared with Hydrogen. If more than one atom of an element enters into the composition of a compound, the number of atoms is shown by a small figure above or below the symbol; as, for instance, in the symbol for water,  $H_2O$ , the index 2 shows that two atoms



of hydrogen are in combination with one of oxygen. A figure placed on the left of a symbol, thus,  $2\text{H}_2\text{O}$ , means two atoms of the compound.

The following is a list of all the known Elementary substances, with their chemical symbols and combining weights. The non-metallic elements are distinguished by capital letters.

Names.	Symbols.	Combining Weights.	Names.	Symbols.	Combining Weights.
Aluminum	Al.	27.5	Molybdenum	Mo.	96
Antimony	Sb.	122	Nickel	Ni.	59
Arsenicum	As.	75	Niobium	Nb.	94.5
Barium	Ba.	137	NITROGEN	N.	14
Bismuth	Bi.	210	Norium	No.	
BORON	B.	10.9	Osmium	Os.	199
BROMINE	Br.	80	OXYGEN	O.	16
Cadmium	Cd.	112	Palladium	Pd.	106.5
Cæsium	Cs.	133	PHOSPHORUS	P.	31
Calcium	Ca.	40	Platinum	Pt.	197.1
CARBON	C.	12	Potassium	K.	39
CHLORINE	Cl.	35.5	Rhodium	Rh.	104.3
Cerium	Ce.	92	Rubidium	Rb.	85.3
Chromium	Cr.	52.5	Ruthenium	Ru.	104.2
Cobalt	Co.	59	SELENIUM	Se.	79.5
Copper	Cu.	63.5	Silver	Ag.	108
Didymium	D.	96	SILICON	Si.	28
Erbium	E.	114.6	Sodium	Na.	23
FLUORINE	F.	19	Strontium	Sr.	87.5
Glucinum	Gl.	9.5	SULPHUR	S.	32
Gold	Au.	196.6	Tantalum	Ta.	182
HYDROGEN	H.	1	Tellurium	Te.	129
Indium	In.	72	Thallium	Tl.	204
IODINE	I.	127	Thorium	Th.	238
Iridium	Ir.	197.1	Tin	Sn.	118
Iron	Fe.	56	Titanium	Ti.	50
Lanthanum	La.	92	Tungsten	W.	184
Lead	Pb.	207	Uranium	U.	120
Lithium	Li.	7	Vanadium	V.	51.3
Magnesium	Mg.	24.3	Yttrium	Y.	61.7
Manganese	Mn.	55	Zinc	Zn.	65
Mercury	Hg.	200	Zirconium	Zr.	89.5

## QUESTIONS ON INTRODUCTION.

1. Define an Element? Distinguish between Organic and Inorganic Chemistry?
2. What is Weight? Explain clearly what is meant by Absolute, Specific, and Atomic Weight?
3. What is the rule for finding the specific gravity (sp. gr.) of a solid? A mass of lead, sp. gr. 11.54, weighs one hundred-weight in air, what will it weigh when immersed in water? (C.)
4. Five grammes of a powdered mineral are put into an empty bottle which holds 100 grms. of water; the bottle is then filled up with water; thus charged it weighs 102.6 grms. Find the sp. gr. of the mineral? Weight of the bottle not taken into account.—*Ans.* 2.08.

## THE NON-METALLIC ELEMENTS.

---

### OXYGEN.

*Symbol O. Combining weight 16. Sp. gr. 1.1056 (Air 1.000).*

This gas was discovered by Dr. Priestley, in England, and Scheele, in Sweden, in the year 1774. It was called "Oxygen" (Acid producer) because at that time all acids were supposed to contain this gas. It is the most widely diffused body in nature: it forms one-fifth of the air by volume, and eight-ninths of water by weight. It is found in all plants and animals, in rocks and earths. When free, it is always observed in the state of gas.

**Preparation.** Some substance, which contains it, is placed in a flask or tube, and heat applied. The substance usually taken is **potassic chlorate** (chlorate of potash). On the application of heat, this soon melts and throws off bubbles of oxygen gas, which may be collected by the displacement of water in inverted glass jars. What remains in the flask is **potassic chloride** (chloride of potassium). The chlorate yields 39.2 per cent. of its weight of oxygen; and 1 oz. gives nearly 2 gallons of gas.

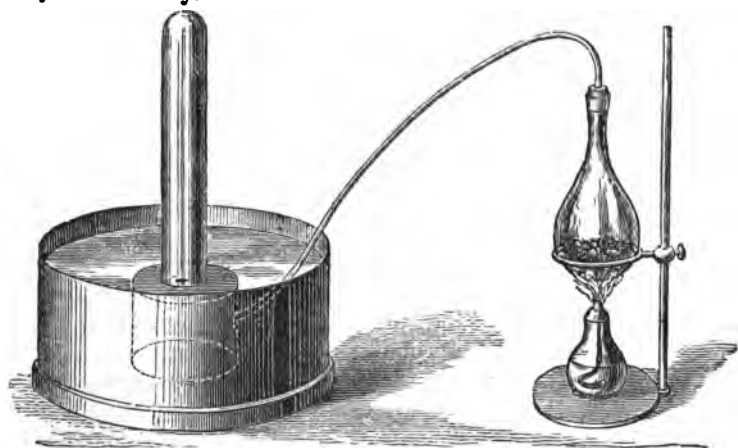
*The change may be shown in symbols thus—*



*The sign + in all chemical equations means "together with."*

Oxygen comes off more quickly when a little powdered **Manganic Peroxide** (black oxide of Manganese) is added to

the potassic chlorate. This gas is also prepared in large quantities by heating the black oxide of **Manganese**, which yields 12·3 per cent. of oxygen by weight. Many other substances yield it readily.



**Properties.** Oxygen weighs a little heavier than atmospheric air. It has no colour, smell, or taste. 100 cub. inches at 60° F. (15°·5 C.) and 30 inches Bar. weigh 34·203 grains. Or one litre at 0° C. and 760 m.m. pressure weighs 1·4336 grammes. It is slightly soluble in water. 100 cub. inches of water at 32° F. (0° C.) take up 4 cub. inches of the gas.

Oxygen is the chief support of animal life: it forms what used to be called the **vital air** of the atmosphere. It has the power of combining with every element except one (Fluorine). One of its most striking properties is its power of supporting combustion. A red hot match plunged into a jar of the gas bursts into flame. Phosphorus burns in it with great brilliancy. Sulphur and Charcoal do the same. Many metals burn vividly in it. A piece of thin iron wire, formed into a spiral, and tipped with sulphur, will, when lighted and plunged into the gas, burn brilliantly. Zinc foil may be used in the same way.

Combustion in air and in oxygen is exactly the same thing. During the process of combustion, the oxygen unites with the burning substance and forms an **oxide**. Thus the burning wire forms with oxygen the **black oxide of iron** ( $\text{Fe}_3\text{O}_4$ ): Phosphorus forms **phosphoric acid** ( $\text{P}_2\text{O}_5$ ): Sulphur forms **sulphurous acid** ( $\text{SO}_2$ ): Charcoal forms **carbonic acid** ( $\text{CO}_2$ ).

These oxides are called compounds, and they weigh exactly the same as the materials used in forming them.

**OXIDES** are very numerous and important. They are conveniently divided into three classes. 1. **Acid oxides**. 2. **Basic oxides**. 3. **Neutral oxides**.

**Acid Oxides** are those which have a sour taste; turn vegetable blue colours to a bright red, and browns to yellow: and combine with Basic oxides to form salts. Sulphuric Acid and Nitric Acid are examples of this class.

**Basic Oxides** have properties the very opposite of Acid Oxides. They have a soapy taste: change vegetable reds to blue or green, and yellows to brown: they neutralize the strongest acids, and combine with acids to form salts. Potass (Oxide of Potassium) and Soda (Oxide of Sodium) are examples of this class.

**Neutral Oxides** are those which have no sour taste, no action upon colouring matter, and do not combine with acids or bases. Water (Oxide of Hydrogen), and Carbonic Oxide are examples of this class.

It frequently happens that a body combines with oxygen in different proportions, and forms a series of oxides. There are, for instance, seven different oxides of the metal Manganese. In order to distinguish them, Greek or Latin numerals are prefixed to the oxide. Thus, if one atom of a metal combines with one atom of oxygen, it is called a **protoxide**; if with two atoms of oxygen, it is called a **binoxide** or **deutoxide**; if with three of oxygen, a **teroxide** or **tritoxide**. The highest oxide is called a **peroxide**. Sometimes two atoms of a metal unite

with three of oxygen, then the name **sesquioxide** is given,—**sesqui** means **one and a half**. Any compound containing less oxygen than a **protoxide** is called a **suboxide**.

When Acid oxides and Basic oxides combine with each other, they form what are called **Salts**. We define a **Salt** to be the union of an **Acid** and a **Base**, or a compound of a pair of oxides, one of which is highly basic, and the other highly acid, according to the explanations given above.

**Salts.** When acids ending in **ic** form salts, this termination is changed into **ate** to express the name of the salt. When sulphuric acid combines with the protoxide of iron, the salt thus formed is called sulphate of iron, or ferrous sulphate ( $\text{Fe}_2\text{SO}_4$ ).

If the name of the acid ends in **ous**, the name of the salt ends in **ite**. Sulphurous acid combines with Baryta and forms the salt sulphite of Baryta, or Baric sulphite ( $\text{BaSO}_3$ ).

The names of some salts end in the syllable **ide**. This termination is used when **two elementary bodies** unite to form a salt. Sulphur and Sodium unite to form the salt Sulphide of Sodium, or Sodic Sulphide.

Care must be taken to distinguish the **ites** from the **ates**, and the **ides** from the **ites**.

Many salts, when crystallized, contain a definite number of atoms of water, and this is called **water of crystallization**.

**Acid Oxides** are distinguished in the following way. The name of the acid, which has the largest proportion of oxygen, ends in the syllable **ic**. An acid containing less oxygen has a name ending in **ous**. Thus in the two compounds,  $\text{H}_2\text{SO}_4$ , and  $\text{H}_2\text{SO}_3$ , the first contains one atom of oxygen more than the second; hence the former is called Sulphuric Acid, and the latter Sulphurous Acid.

**OZONE.**—When electric sparks are sent through the air or dry oxygen, a faint smell is perceived. To the body which produces this smell the name **ozone** is given, from a Greek word, meaning “to give off an odour.” Ozone is supposed to

be oxygen in a more active state after undergoing some change. It possesses bleaching properties, and is present in the atmosphere. It is insoluble in water, and in solutions of the acids and alkalies, potassic iodide (iodide of potassium) excepted.

#### MEASUREMENT OF TEMPERATURE OF BODIES.

**THERMOMETERS.**—Changes in the temperature of bodies are measured by instruments called thermometers. These are closely sealed glass tubes containing Mercury, upon which are marked the freezing and boiling points of water. The distance between these points is graduated according to one of the following scales:—1. Fahrenheit's scale; 2. Centigrade scale; 3. Reaumur's scale. Fahrenheit's scale is the one used in England. According to this, the freezing point of water is marked at 32 degrees, and the boiling point at 212 degrees. In the Centigrade thermometer these points are marked at zero (0 degrees) and 100 degrees respectively; while, according to Reaumur's scale, the same points are marked at zero and 80 degrees respectively. The Centigrade thermometer is used in the greater part of Europe and America, and is now generally employed by scientific men in England. It is very easy to express the scale of one in terms of the other. For example—

Since— $100^{\circ} \text{C.} = 180^{\circ} \text{F.}$

Therefore— $5^{\circ} \text{C.} = 9^{\circ} \text{F.}$

Suppose we want to know what degree of Fahrenheit's scale corresponds to  $60^{\circ} \text{Centigrade}$ . Since  $5^{\circ} \text{C.} = 9^{\circ} \text{F.}$ , therefore  $60^{\circ} \text{C.} = 108^{\circ} \text{F.}$ ; but as Fahrenheit's scale commences with  $32^{\circ}$ , this must be added to  $108^{\circ}$ . Hence the answer will be  $140^{\circ} \text{F.}$  On the same principle, the Fahrenheit may be expressed in terms of the Centigrade. Subtract  $32^{\circ}$ , then, since  $9^{\circ} \text{F.} = 5^{\circ} \text{C.}$ ;  $28^{\circ} \text{F.} = 5 \text{ C.} \times \frac{28}{9} = 15\frac{5}{9}^{\circ} \text{C.}$

The rule of conversion may be expressed thus:—

From C. to F. Multiply by 9, divide by 5, and add 32.

„ F. to C. Subtract 32, multiply by 5, and divide by 9.

## EXPANSION OF GASES BY HEAT.

It has been found, by careful experiments, that gases expand  $\frac{1}{273}$  of their volume at  $0^{\circ}\text{C}.$ , on every increase of  $1^{\circ}\text{C}.$  Thus,

273	volumes of a gas at $0^{\circ}\text{C}.$ , become
274	"      " $1^{\circ}\text{C}.$ , and
275	"      " $2^{\circ}\text{C}.$ , &c.

The fraction  $\frac{1}{273}$  expressed as a decimal, is .003665. If 100 measures of a gas be measured at  $0^{\circ}\text{C}.$ , and then heated to  $1^{\circ}\text{C}.$ , they become 100.3665. Similarly, 100.3665 volumes of gas at  $1^{\circ}\text{C}.$ , become 100 volumes at  $0^{\circ}\text{C}.$  Remembering the fact stated above, if the volume of any gas at  $0^{\circ}\text{C}.$  be given, its volume at any degree of heat and the same atmospheric pressure may be readily found.

## RULE FOR CORRECTION OF GASES FOR TEMPERATURE.

As 273 + given temperature is to 273 + required temperature, so is the given volume to the required volume.

## RELATION OF VOLUME OF GASES TO PRESSURE.

The volume of a gas increases or decreases according to the pressure to which it is subjected. For example, a certain weight of gas occupies 100 cubic inches at the ordinary pressure of the air; if the pressure be doubled, the volume of the gas will be 50 cubic inches, or  $\frac{1}{2}$ ; if the pressure be lessened  $\frac{1}{2}$ , the volume will be 200 cubic inches, or double. This fact of the relation of volume to pressure, is stated in the following law, called Boyle's or Marriotte's Law,—“The volume of a gas is inversely as the pressure.”

From the above statement it is evident that, if the volumes of gases are to be compared, some standard measure of temperature and pressure must be taken. For this purpose it is agreed to take the temperature at  $32^{\circ}\text{F}.$  or  $0^{\circ}\text{C}.$ , and the standard pressure of 30 inches, or 760 millimeters (m.m.) of Mercury.



Suppose it is required to find what weight of Chlorate of Potassium will produce 20 litres of oxygen, at a temperature of 15° C., and barometric pressure of 750 m.m. 20 litres of oxygen at 0° C., and 760 m.m. pressure, will become at 15° C. and 750 m.m., (according to the facts stated above):—

$$20 + 20 \cdot \frac{15}{273} \times \frac{760}{750} = \frac{20 \times (273 + 20) \times 760}{273 \times 750} = 21 \cdot 75 \text{ litres.}$$

Now, since a litre of oxygen at 0° C. and 760 m.m. pressure weighs 1·4336 grms., 20 litres will weigh 28·672 grms.; but 20 litres at 15° C. and 750 m.m. pressure will weigh less, and this is found as follows:—

$$21 \cdot 75 : 20 :: 28 \cdot 672 : 26 \cdot 86 \text{ grms.}$$

Again, since 100 grms. of Chlorate of Potassium yield 39·2 grms. of oxygen, therefore,

$$39 \cdot 2 : 26 \cdot 36 :: 100 : 67 \cdot 24.$$

The weight of salt required is 67·24 grms.

#### QUESTIONS ON OXYGEN, &c.

1. What class of bodies do we call oxides? Name some oxides of an acid nature, and some basic oxides. (L.)
2. Describe and explain the preparation of Oxygen from Potassium Chlorate, (Chlorate of Potassium); and name the substances which are formed when Carbon, Hydrogen, Sulphur, and Phosphorus are respectively burned in an excess of Oxygen. (L.)
3. What weight of Oxygen may be obtained by the decomposition of 53 grains of Chlorate of Potassium? (O.)
4. What is Oxone? Under what circumstances is it formed? What are its properties? (O.)
5. How many grammes of Oxygen can be separated from 100 grammes of Binoxide of Manganese by heat? And how many by the action of Sulphuric Acid? \*
6. Write out Boyle's Law. 20 litres of a gas are measured at 0° C. and 720 m.m. pressure. Required the volume at 760 m.m., the temperature being the same. \*
7. 40 cubic centimetres of a gas are measured at 20° C. How many will it measure at 30° C.? \*

8. 40 cubic centimetres of air are measured at  $15^{\circ}\text{C}$ . and 750 m.m. pressure. Required the volume of air at  $25^{\circ}\text{C}$ . and 800 m.m. pressure. \*

9. A gas measures 117.9 cubic inches at  $9^{\circ}\cdot9\text{C}$ . and 709 m.m. pressure. Required its volume at  $0^{\circ}\text{C}$ . and 1 metre pressure. \*

10. 600 cubic centimetres of gas at  $10^{\circ}\text{C}$ . is warmed until it becomes 730 cubic centimetres. Through what number of degrees  $\text{C}$ . has the gas been heated? \*

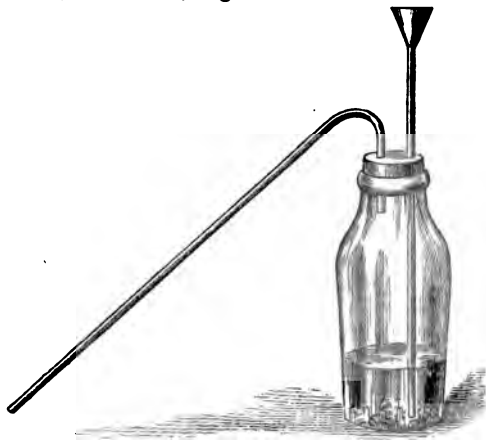
---

## HYDROGEN.

*Symbol H. Combining weight 1. Sp. gr. .0692.*

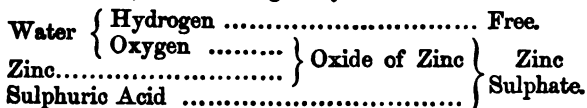
Hydrogen was discovered by Mr. Cavendish, in 1766. It is never found uncombined. It is one of the constituents of water, of which it forms one-ninth part by weight. It is found in all plants and animals. It receives its name from two Greek words, meaning "water producer."

**Preparation.** Hydrogen is obtained by decomposing water. The simplest and easiest way of doing this, is to put some zinc, sulphuric acid, and water, together in a vessel. The water is



immediately resolved into its two elements, hydrogen and oxygen. The oxygen unites with the zinc to form oxide of zinc,

and hydrogen comes off as gas. The oxide of zinc, again, unites with the sulphuric acid and forms the salt, zinc sulphate (sulphate of zinc). The change may be shown thus :—



OR IN SYMBOLS.



One oz., or 15·44 grammes, will disengage 626 cub. in. or about 10 litres of hydrogen from water. Hydrogen may also be obtained from water by passing steam through red-hot iron piping filled with iron-turnings. The oxygen of the steam combines with the heated iron to form black oxide of iron, while its hydrogen passes out at the other end.

The metals **sodium** and **potassium** decompose water. Take a piece of potassium about the size of a pea, wrap it in blotting paper—pass it quickly into an inverted test tube filled with water. The potassium unites with the oxygen of the water to form **potass** (oxide of potassium), and the liberated hydrogen fills the tube.

**Properties.** It is the lightest body in nature, being 16 times lighter than oxygen, and  $14\frac{1}{2}$  times lighter than air. 100 cub. in. of it at 60° F. and 30 in. Bar. weigh 2·14 grains; or in the French metrical system, 1 litre at 0° C. and 760 millimeters Bar. weighs 0·0896 grammes.

Owing to its lightness, it has been often used to fill balloons. It is a bad conductor of sound. A bell rung in a jar of the gas gives a feeble sound. It is slightly soluble in water. 100 volumes of water absorb about 2 volumes of the gas.

It burns in the air when a light is applied to it, with a **pale yellow flame**, giving out much heat, but little light. During combustion, it unites with oxygen, and forms water. If a saucer or glass vessel be held over a burning jet of the gas, it is soon covered with drops of water.

If two volumes of hydrogen be mixed with one volume of oxygen, and a light applied, the gases immediately combine with a loud explosion, and form water. Strong jars are necessary for this experiment. It is well to wrap a cloth round the vessel containing the mixture. On account of the explosive property of hydrogen when mixed with oxygen, care is required in collecting the gas. Time should be allowed for the expulsion of the air from the vessels in use.

If a glass tube, one or two feet in length, be held over a burning jet of the gas, a musical note is produced. This sound is caused by the rapid explosions of hydrogen and oxygen in the confined space of the tube.

Though the flame of this gas has a feeble light, yet, when mixed with oxygen and made to play upon a piece of lime, it produces a most brilliant light, which can be seen nearly 100 miles off on a clear night.

Hydrogen forms numerous compounds with other bodies, but it does not equal oxygen in this respect.



#### COMPOUNDS OF OXYGEN AND HYDROGEN.

There are two compounds of Oxygen and Hydrogen.

1. *Water*. Symbol  $H_2O$ . Combining weight 18.
2. *Peroxide of Hydrogen*. Symbol  $H_2O_2$ . Combining weight 34.

**WATER.** Mr. Cavendish discovered, in 1781, that water consisted of the two gases, hydrogen and oxygen, in the proportion of two atoms of hydrogen to one of oxygen. A claim to the same discovery has been made on behalf of Mr. James Watt. Water may be resolved into its two elements by sending an electric current through it, when mixed with a little sulphuric acid.

Water, when free from admixture, is a clear, colourless, transparent liquid, without taste or smell. At temperatures below 32° F. (0° C.), it freezes, and takes the solid form, as ice. At the temperature 212° F. (100° C.), it boils, and takes the gaseous form, as steam. The boiling point of water depends upon the pressure of the atmosphere. At the level of the sea, the boiling point is 212° F.; but the higher we go, atmospheric pressure becomes less, and therefore water will boil at a less heat than 212°. Heights of places above the level of the sea may be found by noticing the temperature at which water boils.

Water combines with many substances. The compounds thus formed are called **Hydrates**. Bodies that have no water in combination with them are called **Anhydrous** (without water).

Water is sometimes found in the solid state combined with salts: it is then called **Water of Crystallization**. This may be driven off by heat, but, in that case, the crystal crumbles to powder. Sodid Carbonate (Carbonate of Soda), for example, contains Water of Crystallization; but, when the water is driven off, it becomes a white powder.

Water possesses the property of holding many salts in solution. In consequence of this, it is seldom found pure in nature. For the same reason, there are various kinds of waters, such as **Rain water**, **Spring water**, **Mineral water**, **River water**, **Sea water**.

**RAIN Water** absorbs gases floating in the air.

**SPRING Water** varies in composition according to the nature of the soils and rocks through which it passes. It generally holds lime in solution, which gives to the water the character of "hardness."

**MINERAL Waters** contain compounds of iron and sulphur, and are useful for medicinal purposes.

**RIVER Water** contains much vegetable matter and other impurities.

**SEA Water** contains from 3.5 to 4 per cent. of its weight

of salts. These salts are chiefly chlorides or sulphates of sodium, potassium, magnesium, calcium.

Water can be obtained pure by distillation. Distilled water, however, has a sickly flavour, and is unfit for drink. It requires to be charged with air,—“aërated,” to make it palatable: and this is done by filtering it through charcoal.

HYDRIC PEROXIDE, or Peroxide of Hydrogen, ( $H_2O_2$ ) is not found in nature. It differs from water in possessing twice as much oxygen; and because it easily gives off half its oxygen, and turns to water, it is called **oxygenated water**.

It is prepared as follows:—Baryta Peroxide ( $BaO_2$ ) is ground to powder under water, and becomes hydrated ( $BaH_2O_2$ ). When this hydrate is dissolved in hydro chloric acid, baryta chloride ( $Cl_2Ba$ ) is formed, and hydric peroxide ( $H_2O_2$ ).

This latter compound possesses bleaching properties.

#### MEASUREMENT AND WEIGHT OF GASEOUS VOLUMES.

Though the various elements differ in their combining proportions, yet an atomic weight of each element, with the exception of arsenic, phosphorus, and mercury, fills exactly the same space, at the same temperature and pressure. Thus, one atom of hydrogen gas and one of oxygen, of which the combining proportions are 1 and 16 respectively, occupy the same space.

From this fact, if we take a fixed volume of hydrogen as a standard, we shall be able to calculate the absolute weight of a given volume of any gas at the standard temperature and pressure; since the density of a gas is exactly the same as its combining weight. The standard volume-weight unit is 1 litre of pure hydrogen, at  $0^\circ$  C. and 760 m.m. pressure; and the actual weight of this quantity is .0896 grammes. The term **crith** has been given to this standard unit. From this, the weight of a litre of any gas at the same temperature and pressure, is known by multiplying the crith by its combining weight.

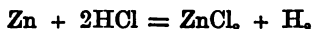
1 litre of oxygen weighs  $16 + 0.0896 = 1.4336$  grammes.  
 1 „ nitrogen „  $14 + 0.0896 = 1.2544$  grammes.

In the case of almost all compound gases, the density is one-half the combining weight. Thus steam, or water-gas ( $H_2O$ ) contains 2 vols. of hydrogen, and 1 vol. of oxygen, and weighs  $2 + 16 = 18$  units; therefore, the single vol. of steam weighs  $\frac{18}{2} = 9$  units; or 1 litre of steam weighs 9 criths  $= 0.8064$  grms.

Since  $0.0896$  grms. of hydrogen,  $0^\circ C.$  and  $760$  m.m. occupy 1 litre  
 $\therefore$  1 gram. „ „ occupies  $11.2$  litres.

Therefore, if we take the atomic weight of an element in grammes, it will occupy  $11.2$  litres, at  $0^\circ C.$  and  $760$  m.m. So  $16$  grms. of oxygen,  $14$  grms. of nitrogen, &c., will occupy this volume. Arsenic and phosphorus occupy one-half this vol., or  $5.6$  litres: mercury, and the molecular weight of a compound, with few exceptions, taken in grms. occupies  $22.4$  litres.

From this we are enabled to find the volume of a substance produced in certain decompositions, or conversely. For example, if it be required to find the number of litres produced, when a kilogramme of zinc is acted on by excess of hydro chloric acid ( $HCl$ ), express the decomposition in the form of an equation, thus—



From the above equation we find that an atomic weight of zinc ( $65.2$ ) gives off a molecular vol. of hydrogen.

Since  $65.2$  grammes of zinc give off  $22.4$  litres of hydrogen.

$\therefore$  1 gramme „  $\frac{22.4}{65.2}$  „

And  $1000$  grammes  $\frac{22.4}{65.2} \times 1000 = 34.35$  litres.

#### QUESTIONS ON HYDROGEN.

- 1.—How much zinc must be weighed out in order to obtain hydrogen enough to fill a gasholder of  $10$  litres capacity? (O.)
- 2.—Describe how you would prepare hydrogen from water without the assistance of an acid; and explain your process. How can a mixture of

two volumes of hydrogen and one volume of oxygen be converted into water? (L.)

3.—State the composition of water by weight, and by volume; and explain how the latter may be shewn. If the elements of water were mixed in a close vessel in the proportion of 1 : 3, and exploded, what would be found in the vessel after explosion? (C.)

4.—How much water would be produced from 28 lbs. of oxygen and 5 lbs. of hydrogen; and would either of the elements be in excess? \*

5.—In what respects does rain-water differ from sea-water? Enumerate the most common constituents of spring-water. How is pure-water obtained from spring-water? (L.)

6.—Water is said to boil at the top of Mont Blanc at 86° C., on Mount St. Bernard at 92° C., and in London at about 100° C. Explain fully the cause of these differences. (L.)

## NITROGEN.

*Symbol N. Combining weight 14. Sp. gr. 0.972.*

This gas is found in a free state in the atmosphere, of which it forms  $\frac{4}{5}$  of the bulk. It is found combined in plants and animals, in coal, and in all nitres. It derives its name, Nitrogen (nitre producer), because it forms the chief part of all nitres. It is also called *azote* (without life), because, though not poisonous, animals cannot live in it. It was discovered by Rutherford in 1772.

**Preparation.** The simplest way of obtaining this gas is to deprive air of its oxygen. For this purpose, burn a fragment of dried phosphorus in a glass vessel inverted in water. The phosphorus should be placed in a floating dish. All the oxygen in the vessel will be removed during combustion, and nitrogen will remain. The phosphorus has united with the oxygen to form phosphoric acid, which dissolves in the water.





Nitrogen may also be prepared by sending a current of chlorine gas through ammonia; also by sending atmospheric air through a tube containing red-hot copper filings. In the latter case, the oxygen of the air unites with the copper to form black oxide of copper, and the nitrogen comes off free.

**Properties.** It has no colour, taste, or smell; it is always in the gaseous state when free. It is a little lighter than air; 100 cub. in. at 60° F. Bar. 30 inches, weigh 30.119 grains; or according to the French system, 1 litre at 0° C. and 766 m.m. pressure weighs 1.2544 grammes. Water dissolves about 2 cub. in. of it in every 100 cub. in. at 32° F.

It neither burns nor supports combustion; it cannot support animal life; its properties are said to be negative.

It forms very powerful compounds with oxygen and hydrogen. With the former it makes the strong acid, nitric acid; with the latter, ammonia.

#### THE ATMOSPHERE.

The atmosphere is a mechanical mixture of several gases, of which oxygen and nitrogen are the most abundant. It is called a mechanical mixture, because the gases in it do not combine to form a chemical compound, but simply mix with each other.

The two chief gases in the air are nitrogen and oxygen; and, from the examination of air in various parts of the world, the proportion in which these two gases are present in 100 parts of air is as follows:—

	By Weight.	By Volume.
Nitrogen.....	76.9	79.1
Oxygen.....	23.1	20.9
	<hr/> 100. <hr/>	<hr/> 100. <hr/>

The relative quantities of these two gases may be shown by the following experiment. Confine air in a glass tube, marked off into 100 equal parts, by inverting it in water: place a piece

of phosphorus in the tube,—the oxygen slowly combines with the phosphorus, and the water ascends 21 parts to take its place, leaving the remaining 79 parts filled with nitrogen.

In addition to the above, the following are also present ;—

**Carbonic Anhydride, (Carbonic Acid)**—from 3 to 6 parts in 10,000.

**Aqueous Vapour**—amount dependent upon the climate.

Traces of Nitric Acid.

Traces of Ammonia.

„ Carburetted Hydrogen.

„ Sulphuretted Hydrogen, and Sulphurous Acid.

The presence of oxygen in air is proved by burning bodies being changed into oxides. Its use in nature, both to animal and vegetable life, has already been mentioned.

The presence of nitrogen is shown in the manner described in the preparation of that gas. It serves to dilute the oxygen of the air, and thus make it suitable for living beings. It also increases the volume of the atmosphere ; and in this way provision is made for winds, and other things useful for man's well-being.

The presence of carbonic acid is shown by exposing a vessel containing lime water in the air. After a while a deposit of chalk (carbonate of lime) will be found in the vessel, in consequence of the union of the carbonic acid of the air with the lime of the water.

This gas supplies plants with one of their chief elements, carbon. In the sunshine, the leaves absorb carbonic gas and decompose it ; they keep the carbon, and give off oxygen. By this wise provision, the air maintains its proper quantity of oxygen.

Nitric acid and ammonia are found in the air combined—as nitrate of ammonia. This descends in rain water, and, entering the roots of plants, supplies them with nitrogen, an element needful for their growth.

Ozone is found in air away from towns, and especially after thunder storms.

A cubic foot of air, at a temperature of 60° F. and 30 inches Bar., weighs 536·96 grains : or a cubic metre of air, at 0° C. and 760 m.m. pressure, weighs 1·2991 kilogrammes. At 60° F. a given volume of air is 816 times lighter than the same bulk of water at the same temperature.

The air extends to a height of 45 miles, and the pressure of a column of air of this height upon a square inch at the level of the sea is 15 lbs. : or, we can say the pressure is equal to a column of mercury 30 inches high, or a column of water nearly 34 feet high. If the density of the air were the same throughout as it is at the sea-level, the height of the atmosphere would then be about 5 miles.

The waters of the sea, rivers, and lakes, hold air in solution ; but this air is richer in oxygen than the atmosphere, because that gas is more soluble than nitrogen. The air dissolved in water contains about 33 parts oxygen in every 100 volumes.

In this way provision is made for the respiration of fishes.

#### COMPOUNDS OF NITROGEN AND OXYGEN.

There are five compounds of Nitrogen and Oxygen, viz.—

1. Nitrous Oxide (laughing gas)— $N_2O$ .
2. Nitric Oxide,  $NO$ .
3. Nitrous Acid or Nitrous Anhydride,  $N_2O_3$ .
4. Nitric Peroxide, or Hyponitric Acid,  $N_2O_4$ .
5. Nitric Acid, or Nitric Anhydride,  $N_2O_5$ .

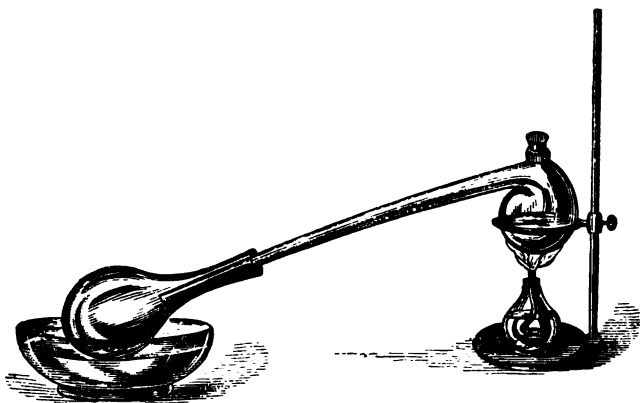
**NITRIC ACID :** or Hydric Nitrate. *Symbol*  $HNO_3$ . *Combining Weight* 63.

This is the most important compound of Nitrogen and Oxygen. In combination with water, hence called **Hydric Nitrate**, it has long been known under the name of **Aqua fortis**. Its true composition was determined in 1785. It may be obtained free from water by exposing dry nitrate of silver to chlorine gas. Nitrate of silver is a combination of nitric acid, silver, and oxygen. The chlorine unites with the silver, forming

chloride of silver, and the oxygen and nitric acid separate. The latter is then called **anhydrous**. Nitric acid in this state was not known till 1849.

Nitric acid is found in the atmosphere, after thunder storms, in the form of nitrate of ammonia; also in tropical climates combined with the metals potassium and sodium. In various parts of India, the soil is crusted over with nitrate of potassium, or potassic nitrate, commonly called saltpetre, or nitre. In the same way, nitrate of sodium occurs in Chili and Peru.

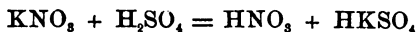
**Preparation.** It is always obtained from one of the nitrates. Place equal weights of nitrate of potassium ( $\text{KNO}_3$ ) and sulphuric acid ( $\text{H}_2\text{SO}_4$ ) in a retort. Apply heat, the potassium unites with the sulphuric acid to form sulphate of potassium,



or dipotassic sulphate, which remains in the retort, while the nitric acid distils over into the receiver. Thus—

Nitrate of	{ Nitric Acid	—————	Hydrated Nitric Acid,
Potassium.			
Sulphuric	{ Potassium	—————	Sulphate of Potassium.
	Acid		
	{ Water	—————	
	{ Sulphuric Acid		

OR IN SYMBOLS.



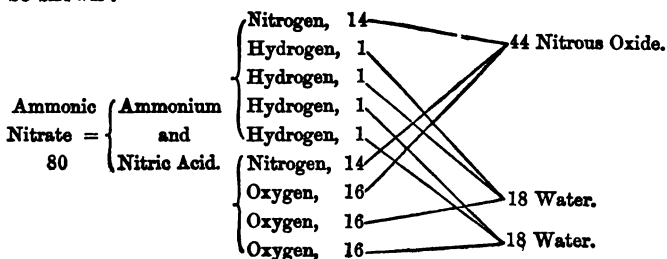
Saltpetre and Sulphuric Acid give Nitric Acid and Sulphate of Potassium.

**Properties.** Nitric acid, when pure, is colourless, but it is usually of a yellowish colour, on account of the presence of some of the lower oxides of nitrogen. Its sp. gr. at 60° F. (15°C.) is 1.517 : it boils at 184° F., and in every 100 parts there are 85.72 of nitric anhydride ( $N_2O_5$ ) and 14.28 of water.

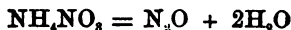
It is very corrosive, and dissolves nearly all the metals except platinum and gold. It destroys all animal textures, and stains the skin, wool, and feathers of a bright yellow colour. If copper, zinc, tin, or iron filings be thrown into nitric acid, ruddy fumes (hyponitric acid) are given off, and oxides of the metals are formed, which combine with the acid, and form nitrates.

Most of the salts called nitrates are soluble in water, fuse readily when heated, and at a high temperature are decomposed.

**NITROUS OXIDE**, or Protoxide of Nitrogen : symbol  $N_2O$  : combining weight 44.—is prepared by heating ammoniac nitrate (nitrate of ammonium) in a glass retort, when the salt is resolved into nitrous oxide and water. The change may thus be shown :



OR IN SYMBOLS.



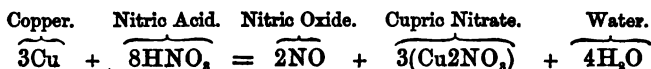
**Properties.** It is without colour and smell, but possesses a sweetish taste. Water at 60° F. dissolves about two-thirds its volume. It becomes a liquid when exposed either to great

pressure or intense cold. Its sp. gr. is 1.527. It is, like oxygen, a brilliant supporter of combustion, but it is distinguished from oxygen by its solubility in water, its taste, and giving no ruddy fumes when mixed with nitric oxide. If respired for a few minutes it produces an intoxicating effect, hence it has the name of **laughing gas**.

**NITRIC OXIDE**, or Binoxide of Nitrogen. *Symbol* NO. *Combining weight* 30. *Sp. gr.* 1.039.

This gas is prepared by pouring diluted nitric acid upon copper clippings in a retort; it may be collected over cold water. One ounce or 30 grms. of copper, acting upon 4 ounces of diluted nitric acid, yields nearly 420 cubic inches, or about 7 litres of nitric oxide.

During this decomposition, only one-fourth of the nitric acid is changed into nitric oxide. 2 out of every 8 parts of the acid give up the oxygen to 3 of copper, forming 3 atoms of protoxide of copper, 1 atom of water, and 2 of nitric oxide. The 3 atoms of the oxide of copper unite with the remaining 6 of nitric acid, and form 3 atoms of cupric nitrate (nitrate of copper), and 3 of water are also separated. Thus, in symbols:—



**Properties.** It is colourless, incombustible, irrespirable. It has never been liquefied; it has a strong attraction for oxygen. If a jar of the gas be exposed to the air, dense ruddy fumes are produced, in consequence of the union of the nitric oxide with the oxygen of the air, forming hyponitric acid ( $\text{N}_2\text{O}_4$ ). Many burning bodies are extinguished when plunged into it, but phosphorus burns vigorously in it.

**NITROUS ACID**, or Nitrous Anhydride. *Symbol*  $\text{N}_2\text{O}_3$ . *Combining weight* 76.

This is prepared by mixing 4 volumes of nitric oxide and 1

of oxygen in a flask, both in a dried state, and then freezing the mixture, which condenses into a blue volatile liquid.

Water decomposes it. Nitrous Acid forms salts with the metals of the alkalies, which are called **nitrites**.

**NITRIC PEROXIDE**, or Hyponitric Acid. *Symbol*  $N_2O_4$ . *Combining weight* 92.

This is prepared by mixing 2 volumes of nitric oxide with 1 of oxygen, or by heating dry nitrate of lead in a retort. The dense fumes which nitric oxide forms when exposed to the air are chiefly the vapour of hyponitric acid. It can be condensed by cold into a volatile liquid, yellowish red at ordinary temperatures, but colourless below zero.

#### QUESTIONS ON NITROGEN, &c.

1. One hundred volumes of Nitrogen are given. How many volumes of Oxygen must be added in order to obtain a mixture of the same composition as pure atmospheric air? (L.)
2. State the constitution of the Atmosphere, and shew how the proportions of its constituents may be determined. How is it affected by fires, and by the action of animal life? What is the natural compensation for this? (C.)
3. Name the most important natural substances in which Nitrogen is contained; and describe how you would prepare Nitrogen. Enumerate the principal properties of Nitrogen, and show how it may be distinguished from Oxygen, Hydrogen, and Carbonic Acid. (L.)
4. Explain how you would prove the pressure of Water, Carbonic Acid, Oxygen, and Nitrogen in atmospheric air. (L.)

#### QUESTIONS ON COMPOUNDS OF NITROGEN AND OXYGEN.

1. What is the action of Nitric Acid upon the metals—Iron, Copper, Zinc, and Tin? (L.)
2. Describe the mode of preparation of Nitric Acid, expressing in Symbols the reactions which take place. Distinguish between a Nitrate and a Nitric. (C.)
3. A cubic inch of Nitrous Oxide weighs .475 grains. What quantity of Nitrate of Ammonia is required to produce 5 cubic feet of this gas? (O.)

4. Describe the combinations of Oxygen with Nitrogen, and the processes by which they are respectively obtained. (O.)
5. What is the composition of Laughing Gas? How is it prepared? (C.)
6. What is the result of mixing over water Nitric Oxide and Oxygen? (O.)
7. Describe in words, and represent in symbols, the effects of heat upon Nitrate of Ammonia. Give some account of the chemical properties of the gaseous product. (L.)

---

## CARBON.

*Symbol C. Combining weight 12.*

This is the first solid we have to mention. It is never found free in the state of a liquid or a gas. It is found uncombined in the solid state, as **diamond** and **plumbago**.

**Charcoal** is another form of this element. Though these substances are all carbon, yet they differ very remarkably in their properties.

Carbon is found in all animals and vegetables. It is found in the air, as carbonic acid. In combination with lime and magnesia it forms a great part of the earth's crust, as limestone, chalk, marble, corals, shells, &c.

**Diamond.** Sp. gr. 3.33 to 3.55. Lavoisier discovered this substance to be pure carbon in 1775-6, by burning it in oxygen, and converting it all into carbonic acid. It is the hardest substance known; and, when cut, possesses a brilliant lustre, and power of refracting light. It is found in a crystallized form in certain rocks and gravels. The most famous diamond mines are those of Golconda and Bundelcund in India, of Borneo, and of Brazil. It cannot be melted at any heat; but, placed between the poles of a powerful voltaic battery, it becomes a black mass resembling coke.

**Plumbago, or Graphite,** commonly called **Black-lead**. Sp. gr. 2.15 to 2.35. It occurs in veins in the granitic and primitive rocks, and is found in Cumberland, the United States, Ceylon, and Siberia. It is generally found mixed with a little iron and clay. The best variety is used for lead-pencils. The



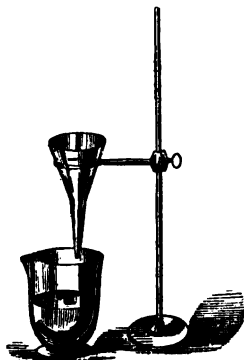
coarsest varieties are used for coating iron, polishing grates, and making crucibles to stand great heat.

**Charcoal** is an artificial variety of carbon ; so also are the varieties known as **coke**, **lamp-black**, and **animal charcoal**. Charcoal is formed by heating wood to redness in closed retorts, until all the water, oils, and gases are driven off. It is used in the manufacture of gunpowder, and in chemical experiments. It is very porous ; it absorbs about ninety times its own volume of ammonia gas, and nine times its volume of oxygen. It destroys bad odours, and is used for this purpose in hospitals, &c.

**Coke** is the name given to the residue of coal, when treated in the same way as wood in the manufacture of charcoal. It is used as fuel, and especially for locomotive engines. It gives off little or no smoke, produces a steady and intense heat, and does not choke the draught by swelling or caking.

**Lamp-black** is the soot of imperfectly burnt combustibles. Mixed with linseed-oil and soap, it forms printers' ink. It is the purest form of artificial carbon.

**Animal charcoal** is obtained by heating bones in a retort.



It is also known as **bone-black**, and **ivory-black**. On account of its great bleaching power, it is much used by sugar refiners to remove the colouring matter in raw sugar. This property may be seen by placing a little of it in water, coloured with indigo or litmus ; boil the liquid, and then filter it, when a colourless liquid will be obtained.

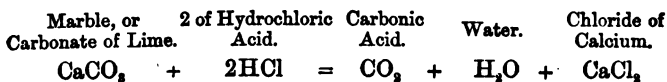
Carbon is a very unalterable substance. Poles driven into the earth, and piles driven into the beds of rivers, are first charred to preserve them from decay. Carbon, when heated, unites with oxygen, forming two compounds—carbonic oxide and carbonic acid.

## COMPOUNDS OF CARBON AND OXYGEN.

**CARBONIC ACID.** (Carbonic Anhydride.) *Symbol*  $\text{CO}_2$ : *Combining weight* 44. *Sp. gr.* 1.529.

This gas is found free in the atmosphere, in the water of many mineral springs, and, combined with lime and magnesia, in limestone, corals, marble, shells, chalk, &c.

**Preparation.** It may be prepared in two ways, known as the analytic and synthetic processes. The synthetic mode is to burn carbon in oxygen, when each atom of carbon unites with two atoms of oxygen, and forms carbonic acid. It is obtained **analytically** by pouring some strong acid, as hydrochloric acid, on pieces of marble. Marble is a compound of lime ( $\text{CaO}$ ) and carbonic acid, and gives off the latter when hydrochloric acid ( $\text{HCl}$ ) is poured upon it. The chemical changes may thus be shown :—



**Properties.** It is a colourless, invisible gas, with a faintly acid smell and taste. Water dissolves about its own volume of the gas; but the gas becomes more soluble if pressure be applied. When the pressure is removed the gas escapes with effervescence. Advantage is taken of this property to make **soda-water**, as it is called. This gas is half as heavy again as air, and may be poured from one vessel into another. Under a pressure of  $35\frac{1}{2}$  atmospheres at  $32^{\circ}$  F., it is liquefied, and may be frozen into a snow-white solid. It cannot support combustion or animal life. A lighted candle put into a jar of the gas immediately goes out. If it be present in the air more than 3 or 4 per cent. it destroys animal life, hence the danger of crowded and ill-ventilated rooms.

It faintly reddens moist litmus paper; but when the gas goes off, the blue colour again returns. By combining with bases, it forms a numerous and important class of salts, called **carbonates**.

**Natural Sources of Carbonic Acid.** (1). It is a product of animal respiration. If air from the lungs be sent through lime water, the latter becomes milky, on account of the formation of chalk. (Carbonate of lime.)

(2). It is produced during the process of fermentation. In breweries many accidents have happened by people descending into the vats.

(3). It is a product of burning limestone in lime-kilns. In volcanic districts it often escapes from openings in the earth. The Dog's Grotto near Naples is an example of this.

(4). It is a product of the combustion of substances in air. Thus coal, gas, tallow, &c., when burning, give off this gas abundantly. Two tallow candles, when burning, produce as much carbonic acid as a man gives off in respiration in the same interval of time.

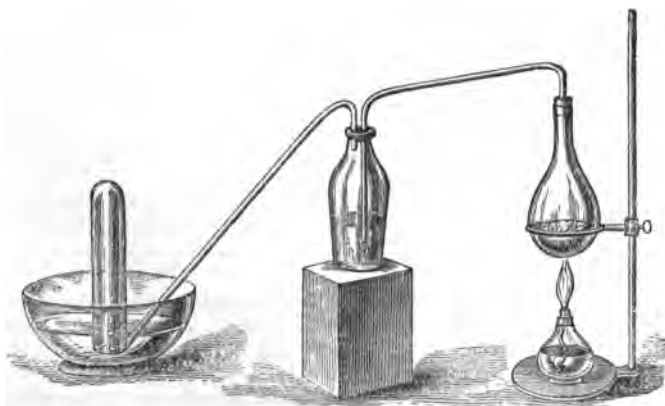
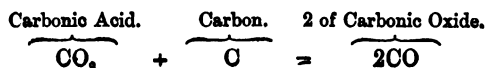
(5). It is found in coal-mines after explosions of **fire-damp**, and is called by miners **choke-damp**. It accumulates in wells

and pits; hence it is necessary to let down a lighted body before descending into one of these.

**CARBONIC OXIDE.** *Symbol* CO. *Combining weight* 28. *Sp. gr.* 0.967.

**Preparation.** This gas contains one atom less of oxygen than carbonic acid. It may be obtained by passing carbonic acid through a red-hot iron tube, when one atom of oxygen unites with the iron, and carbonic oxide comes out at the other end. In this way, one atom of carbonic acid produces one of carbonic oxide.

It may also be produced by sending carbonic acid through a porcelain tube containing red-hot charcoal. In this case one atom of carbonic acid produces two of carbonic oxide: for as the acid passes over the red-hot carbon, it parts with half its oxygen to the carbon, thus producing an atom of carbonic oxide. In symbols—



This latter process is constantly taking place in ordinary fires. The burning coals at the bottom of the grate produce

carbonic acid, which, passing through the red-hot part in the centre, takes up carbon, and becomes carbonic oxide; this burns at the top of the fire, and changes again to carbonic acid.

For purposes of experiment, this gas may be conveniently prepared by heating together the deadly poison, oxalic acid, and strong sulphuric acid. The gases, carbonic acid and carbonic oxide, come off, but these may be separated by sending them through a solution of potass.

**Properties.** It is a transparent, colourless gas: sparingly soluble in water: it is very poisonous,—very small quantities, if taken into the lungs, cause death. The fatal effects of charcoal fires and limekilns are due to the presence of this gas. It does not support combustion, but burns with a pale blue light, producing carbonic acid.

#### QUESTIONS ON CARBON AND COMPOUNDS WITH OXYGEN.

1. What weight of Carbon is needed for the formation of 88 grammes of Carbonic Acid? What is the volume of this weight of acid at 0°C. and 30 inches bar.? (L.)

2. Describe how you would prepare Carbonic Acid from chalk; and explain your process by means of symbols. How would you test whether the air in a well consists of ordinary air or of Carbonic Acid? (L.)

3. How much Carbon could be completely burned by a ton of air? (L.)

4. What volume of air is required for the combustion of one gramme of Carbon, the temperature being 15°C. and the height of the barometer 722 m.m.? (O.)

5. How many litres of Carbonic Acid (measured at 0°C. and 760 m.m.) can be converted into Carbonic Oxide by 12 grammes of Carbon? and how many litres of Carbonic Oxide, at the same temperature and pressure, are thus produced? \*

6. Eighteen cub. inches of Carbonic Oxide exploded, with 9 cub. inches of Oxygen, yield 18 cub. inches of Carbonic Acid. The specific gravities of Oxygen and Carbonic Acid are 1.1055 and 1.5197: find the sp. gr. of Carbonic Oxide. (O.)

7. A current of pure air passes through a bright coke fire. What gases are contained in the air beyond the fire? Give a short description of each of them. (L.)

## COMPOUNDS OF CARBON AND HYDROGEN.

These compounds are very numerous, but their discussion belongs to the division of chemistry called **organic**. There are, however, two important compounds, called **light** and **heavy carburetted hydrogen**, which must be considered here.

**LIGHT CARBURETTED HYDROGEN.** *Symbol*  $\text{CH}_4$ . *Combining weight* 16. *Sp. gr.* 0.5576.

This gas is produced in stagnant pools by decaying vegetable matter, and is for that reason called **marsh gas**. It is given off from seams of coal in mines under the name of **fire damp**; and it is one of the chief constituents of coal gas.

It may be obtained by heating together 40 parts of crystallized acetate of sodium, 40 parts of solid caustic potass, and 60 parts of powdered quicklime.

**Properties.** It is a colourless, tasteless, inodorous gas. Next to hydrogen, it is the lightest known substance. It burns with a yellow flame, forming carbonic acid and water. If mixed with twice its volume of oxygen, or ten times its volume of air, a violent explosion takes place on the application of a light. This is the mixture which causes explosions in coal-mines. The gases formed after the explosion, viz.—carbonic acid, nitrogen, and steam, are commonly named the “**after damp**.” Sir Humphry Davy and George Stephenson invented lamps to prevent these fatal explosions in coal-mines.

**HEAVY CARBURETTED HYDROGEN.** *Symbol*  $\text{C}_2\text{H}_6$ . *Combining weight* 28. *Sp. gr.* 0.967.

This gas is an important constituent of coal-gas. It is known by the name of **Olefiant Gas**, because it forms an oily liquid when mixed with an equal volume of chlorine gas.

**Preparation.** Mix 1 measure of alcohol (spirits of wine)

with 3 measures of sulphuric acid; heat the mixture in a retort, and olefiant gas comes off with water. Thus:



**Properties.** It is a transparent, colourless gas, having a sweetish taste. Under great pressure it may be liquefied. It is nearly double the weight of Marsh-gas, hence it is called **heavy carburetted hydrogen**. It does not support life or combustion; but it burns with a white luminous flame.

When mixed with an equal volume of chlorine, oily drops are formed, hence the name of **olefiant** (oil producing) gas. This oily body is also called **Dutch liquid**, because it was discovered in Holland. If mixed with 2 volumes of chlorine, it burns with much smoke, because its carbon is rejected, whilst its hydrogen unites with the chlorine to form hydrochloric acid.

When mixed with 3 volumes of oxygen, it explodes on the application of a light.

**COAL-GAS.** Coal-gas is a mixture of several gases, obtained by heating coal in closed iron retorts. The best coal for this purpose is **cannel coal**. When the coal is heated it gives off several substances, such as **gas, tar, ammonia, water**. As these pass through a series of iron pipes, the tar, ammonia, and water are deposited, and the gas, after further purification, is received in large cylindrical vessels suspended in water.

Coal-gas is purified in the following way:—As it issues from the retort, it passes through a large iron chamber half full of liquid into large perpendicular iron pipes, and there, in cooling, deposits the tar and ammoniacal liquor. The gas then passes into a chamber containing lime, which removes the carbonic acid and hydrosulphuric acid, and thence into the reservoirs or gas holders.

The gas which passes into the receiver is a mixture of several gases. The following are usually present :—

Light carburetted hydrogen—about 40 or 50 per cent.

Hydrogen— ditto.

Olefiant gas— about 5 per cent.

Carbonic oxide.

Nitrogen.

Certain compounds of carbon and hydrogen.

Sulphuretted hydrogen.

The proportion of the above substances depends upon the kind of coal employed. Olefiant gas is the chief illuminating substance. In comparing the light of coal-gas with that of candles, the gas from cannel coal, burning at the rate of five cub. ft. per hour, is said to be equal to 34·4 candles, burning 120 grains per hour, and common coal-gas equal to 13 candles.

#### THE NATURE AND STRUCTURE OF FLAME.

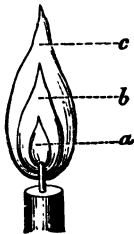
Flame is **gas** heated to such an extent that it becomes luminous. In order that flame may give off much light, it must contain solid matter, which becomes heated to whiteness. The amount of light does not depend upon the degree of heat, for the hottest flames, as the oxyhydrogen flame, give little light.

The flame of a candle, or lamp, has **three** distinct parts :—

(a) The dark part around the wick, containing the unburnt gases, called the **area of no combustion**.

(b) The luminous part about (a), in which the hydrogen and carbon from the tallow are only partially burnt, called the **area of partial combustion**.

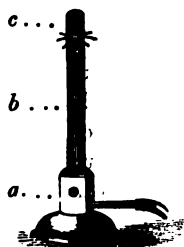
(c) The outside cone about (b), in which





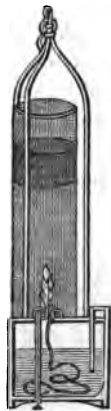
complete combustion takes place, giving off great heat, but little light: this is called the area of complete combustion.

In (a) the gases do not burn, because the oxygen of the air cannot reach that cone. In (b) the carbon in the solid state is heated to whiteness, because sufficient oxygen does not reach that part to burn it, hence the great light of this cone. In (c) the supply of oxygen is sufficient to burn the carbon, converting it into carbonic acid, and making the flame non-luminous.



By means of the **Bunsen lamp**, a flame of complete combustion may be obtained. This lamp consists of an ordinary gas-burner, covered over with a tube about 4 inches long, open at the bottom. The air passes in at the bottom (a), mixes with the gas in the tube (b), and burns on the top at (c) with a blue smokeless flame. This lamp is in general use in laboratories.

#### THE DAVY LAMP.



This lamp was invented by Sir Humphry Davy to protect miners from explosions in coal-mines. It consists of an oil lamp, enclosed in a cylinder of fine wire gauze, and furnished with the means of trimming the wick from the outside. The principle of the lamp depends upon the following facts:—Great heat is required to produce a flame. If the temperature of burning gas be lowered beyond a certain point, the flame goes out. A coil of copper wire, for example, placed over a flame, conducts away so much heat as to extinguish it. The flame, therefore, within the Davy lamp cannot pass through the wire gauze, because the metal reduces the temperature to such a point as to put out the flame.

This simple contrivance preserves mines from explosions of fire-damp. Marsh-gas, as we have seen, is combustible; but when mixed with a certain quantity of air, it becomes explosive. The gas in the pits passes through the wire gauze of the lamp, and burns within; but no flame can pass without to explode the mixture. Marsh-gas does not become explosive if mixed with less than 3 times its volume of air, or more than 18 times its bulk.

### THE BLOW-PIPE.

This is a small instrument for increasing the activity of combustion by throwing a jet of air into the interior of a flame. The blow-pipe flame is divided into two parts, the outer (*b*) is called the **oxidating** flame, and the inner (*a*) the reducing flame. In the inner flame the unburnt combustible gases, at a very high temperature, unite with the oxygen of any metallic oxide placed in it, and thus reduce the oxide to the metallic state, hence it is called the **reducing** flame. On the other hand, the heated air at the point (*b*) of the outer flame, imparts oxygen to a metal held there, and turns it into an oxide, hence it is called the **oxidating** flame. These properties of the blow-pipe flame may be seen by exposing a little lead on charcoal to the outer flame, when it becomes the oxide of lead, and melts into a transparent globule. If this be now put into the inner flame, the oxygen is removed, and the lead appears in the metallic state.

The blow-pipe is very useful for examining substances. Its use requires a continual blast, which requires some practice to learn.



## COMPOUNDS OF CARBON AND HYDROGEN, AND COAL-GAS.

1. How much pure air ought to be mixed with 100 measures of Fire-damp (Marsh-gas) in a coal-mine, in order to supply material for the complete combustion of the Marsh-gas during explosion? (L.)
2. A cubic foot of Marsh-gas is mixed with 2 cubic feet of Oxygen, and exploded. What will be the resulting compounds, and what space will they occupy? (C.)
3. Explain the action of the Miners' Safety Lamp. (C.)
4. What chemical properties are characteristic of Olefiant gas? (O.)
5. What is the cause of an explosion in coal mines, and whence arises the danger of descending immediately after an explosion? (O.)
6. What volume of Olefiant gas must be used to form 10 cubic centimetres of Dutch liquid, of which the sp. gr. is 1.25? (O.)
7. Name the most important constituents, and some of the commonest impurities of Coal-gas. Give their chemical formulas, and explain what are the products to which they severally give rise when the gas is burnt. State which of the impurities may be removed, and by what means. (C.)
8. To what constituent does common Coal-gas chiefly owe its illuminating power? State how many volumes of Oxygen are necessary to its complete combustion, and name the resulting product. (C.)

## FLAME AND THE BLOW-PIPE.

1. Explain the nature of the chemical changes effected by the inner and outer flames of the Blow-pipe. (C.)
2. Describe the structure of the flame of a common candle, and explain why the flame loses its brightness when a current of common air is blown into it. (L.)
3. Explain fully why the flame of a common gas-light is brighter than the flame from a Bunsen burner. (L.)

---

CHLORINE.

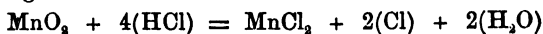
*Symbol Cl. Combining weight 35.5. Sp. gr. 2.47.*

The substances chlorine, bromine, iodine, and fluorine are very much alike in their properties. They have great affinity

for hydrogen, but very little for oxygen. They combine with metals to form salts, and in this respect differ from all the other elements; as, for example, chlorine and sodium form common salt. For this reason they are called **simple salt radicals**.

Chlorine, the most important of the salt radicals, was discovered by Scheele in 1774. It is found combined in nature as chloride of sodium, or common salt; it is found in soils and natural waters.

**Preparation.** It may be obtained from any chloride. The most convenient way is to heat in a retort the black oxide of manganese ( $\text{MnO}_2$ ), and hydrochloric acid ( $\text{HCl}$ ). Four atoms of hydrochloric acid act upon one atom of black oxide of manganese; two atoms of chlorine are given off as gas, and there remain in the retort two atoms of water, and one of chloride of manganese. Thus:—



The gas may be collected over **warm** water, because of its solubility in **cold** water, or by displacement of air.

**Properties.** It is a gas of a greenish yellow colour, hence its name chlorine, from the Greek word  $\chi\lambda\omega\rho\delta\varsigma$ —green. It has a strong, suffocating odour, but when present in small quantities it has the smell of sea-weed. It is heavier than air; 100 cubic inches at the ordinary temperature and pressure weigh 77 or 78 grains. Under the pressure of 4 atmospheres it becomes a liquid. One volume of water at 60° F. dissolves 2.37 volumes of chlorine.

It has great affinity for metals. Powdered metallic antimony showered into a jar of the gas catches fire. Other powdered metals and phosphorus will also burn when treated in the same way. The salts formed by the union of chlorine with the metals are called **chlorides**.

Chlorine, when associated with water, has strong bleaching power. Most of the vegetable colouring matters contain

hydrogen, and since chlorine has great affinity for this gas, the colouring matter is decomposed. This property may be shown by the action of chlorine upon writing ink. The gas in the free state, and in the form of bleaching powder, is used in cotton, linen, and paper manufactures.

It is also useful as a disinfectant. It is used for this purpose in hospitals, &c.

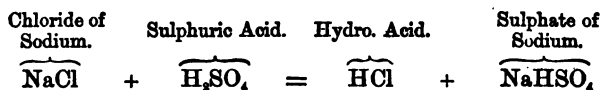
**HYDROCHLORIC ACID.** *Symbol* HCl. *Combining weight* 36.5. *Sp. gr.* 1.261.

This most important compound of chlorine and hydrogen is also called **muriatic acid**, **marine acid**, **hydric chloride** (chloride of hydrogen).

**Preparation.** It may be prepared in two ways—

(1). By mixing equal volumes of chlorine and hydrogen, and exposing them to sunlight, or applying a lighted match, when they will unite with explosion, and form hydrochloric acid.

(2). By heating in a retort common salt (chloride of sodium) and sulphuric acid. Hydrochloric acid comes off as gas, which may be collected by displacement of air, and hydric-sodium sulphate (sulphate of sodium) is left behind in the retort. Thus :—



One ounce, or 30 grammes of common salt, yields 700 cubic inches, or 12 litres of the gas.

**Properties.** It is a colourless, invisible gas, 1.269 times heavier than air; it possesses a most pungent odour, and very soluble in water. It may be liquefied under the pressure of 40 atmospheres. When allowed to escape into the air, it produces white fumes by combination with the moisture of the air. It is neither combustible nor a supporter of combustion.

Commercial hydrochloric acid is a solution of the gas in water. Water absorbs about 480 times its bulk of the gas. The strongest commercial hydrochloric acid contains about 43 per cent. of the gas. The commercial acid is generally impure; its yellow colour is owing to the presence of iron, arsenic, sulphurous acid, &c.

When this acid comes in contact with a metallic oxide, it does not form a salt like an oxygen acid, as sulphuric acid does. If, for example, the oxide of sodium and hydrochloric acid come together, the hydrogen of the acid unites with the oxygen of the oxide to form water, and the chlorine unites with the metal to form the salt—chloride of sodium. But if sulphuric acid come in contact with the same oxide, the two unite and form the salt—sulphate of sodium.

The test for hydrochloric acid is nitrate of silver. This gives a white precipitate in the free acid and its compounds, insoluble in nitric acid, but soluble in ammonia.

Two parts of hydrochloric acid, mixed with one of nitric acid, form *aqua regia*, which dissolves gold, platinum, and other substances.

#### COMPOUNDS OF CHLORINE AND OXYGEN.

These gases do not combine directly with each other, but they may be formed indirectly into compounds. Three of these are obtained in a separate state, and four in combination with water.

		Symbol.		Combining weight.
Hypochlorous Anhydride		$\text{Cl}_2\text{O}$	... ..	87
Chlorous Anhydride	...	$\text{Cl}_2\text{O}_3$	... ..	119
Chloric Peroxide	...	$\text{ClO}_3$	... ..	67·5
Hypochlorous Acid	...	$\text{HClO}$	... ..	52·5
Chlorous Acid	...	$\text{HClO}_2$	... ..	68·5
Chloric Acid	...	$\text{HClO}_3$	... ..	84·5
Perchloric Acid	...	$\text{HClO}_4$	... ..	100·5

**HYPOCHLOROUS ANHYDRIDE** is obtained by passing chlorine through a tube containing red oxide of mercury. It is a colourless gas, but it may be reduced to a deep red liquid, which sends off a suffocating dark vapour. This vapour is very explosive. A solution of this gas and hypochlorous acid has powerful bleaching properties.

**CHLOROUS ANHYDRIDE** is a yellowish green gas, dangerous to prepare on account of its explosive properties. It dissolves in water, and forms chlorous acid.

**CHLORIC PEROXIDE** is a dark coloured explosive gas, and is obtained by the action of sulphuric acid on chlorate of potassium. By means of this gas phosphorus may be made to burn under water. Put a few drops of phosphorus in a long narrow glass, add some chlorate of potassium, and cover them with water, then add a few drops of sulphuric acid by means of a glass funnel, and the phosphorus will take fire.

**CHLORIC ACID** is only of interest as combining with metallic oxides to form salts called **chlorates**. All chlorates, when heated, part with their oxygen, and they are all soluble in water.

**PERCHLORIC ACID** is obtained by heating chlorate of potassium until one-third its oxygen is driven off. The mass which remains is a chloride and a chlorate, and from these the perchlorate of potassium is formed. If strong sulphuric acid be mixed with the dry perchlorate, perchloric acid comes off, and condenses in the receiver as a colourless fuming liquid. It is one of the most powerful oxidizing agents known; a drop allowed to fall on paper, wood, or charcoal, produces combustion with violent explosion.

Perchloric acid forms the salts called **perchlorates**, which are all decomposed by heat. They may be distinguished from the

chlorates by not giving off a yellow gas when moistened with sulphuric acid.

### COMPOUNDS OF CHLORINE AND NITROGEN.

These two gases may be made to combine indirectly by sending chlorine gas through a solution of ammonia, when drops of an oily liquid are seen to form. This liquid is known as **chloride of nitrogen** ( $\text{NCl}_3$ ). It explodes with fearful violence when touched, producing free nitrogen and chlorine.

### COMPOUNDS OF CHLORINE AND PHOSPHORUS.

1. **Terchloride of phosphorus** ( $\text{PCl}_3$ ), a clear volatile liquid, obtained by heating phosphorus and corrosive sublimate.

2. **Perchloride of phosphorus** ( $\text{PCl}_5$ ), a snow-white solid, obtained by placing phosphorus in a vessel of dry chlorine gas.

### QUESTIONS ON CHLORINE AND ITS COMPOUNDS.

1. Describe how you would prepare Chlorine, and explain your process. Also give an outline of the chief chemical and physical properties of this element. (L.)

2. Explain the action of Chlorine as a bleaching agent. A litre of Hydrogen gas at  $0^\circ \text{C}$ . and 760 m.m. pressure weighs 0.08936 gramme. What is the weight of the same quantity of Chlorine gas under the same conditions? (C.)

3. Describe and explain the process of making Hydrochloric Acid. (C.)

4. Determine the proportions in which the substances should be employed in making Hydrochloric Acid, so that there may be no waste; having given the equivalents of Sodium, Sulphur, and Chlorine to be as 23 : 16 : 35.5. (C.)

5. What weight of Oxygen, Nitrogen, and Chlorine, respectively, will combine with one part by weight of Hydrogen? And what volume of each of these gases will combine with one volume of Hydrogen? (O.)

6. To a solution of Hydrochloric Acid an excess of Carbonate of Sodium was added; the liquid was evaporated, and the dry residue weighed. The weight of Carbonate of Sodium added was 1.5 grms., that of the residue 1.6507 grms. What weight of Hydrochloric Acid did the solution contain? (O.)



7. A piece of marble is placed in dilute Hydrochloric Acid; state the effects observed, and the compounds formed. If a piece of iron be used instead of marble, state the results. (C.)

8. If, in the case of question 7, 4 fluid ounces of the acid will dissolve 1000 grains of marble, what quantity of iron would you expect an equal quantity of the acid to dissolve; the equivalents of calcium and iron being 20 and 28 respectively. (C.)

9. What is Chloride of Lime? Explain how it is prepared, and how it is applied in bleaching cotton fabrics. Would you expect stains of lamp-black or of nitrate of silver to be removed by these means? Why so? (C.)

## BROMINE.

*Symbol Br. Combining weight 80. Sp. gr. as a gas 5.54 : as liquid 3.187.*

This element is much like chlorine in its properties and compounds. It was discovered by Balard, in 1826, in *bittern*,—the mother liquor of sea-water. It was named Bromine from *βρῶμος* (Bromos)—a *stench*, on account of its disagreeable smell.

It does not occur free in nature, but it is found combined with sodium and magnesium in sea-water and mineral springs.

**Preparation.** Sea-water is evaporated until the more insoluble salts, which it contains, are removed. The liquor is then called *mother liquor*, or *bittern*. Chlorine gas is sent through this, and liberates bromine from those metals with which it is combined. The Bromine dissolves in the liquid, and gives it a yellow colour; this, agitated with ether, parts with its bromine, which dissolves in the ether.

Caustic potash, added to this solution, forms a bromide and a bromate of potassium; the ether may be evaporated and the bromate separated by heat. The remaining bromide, mixed with black oxide of manganese, and dilute sulphuric acid, and heated, gives off bromine as a deep red vapour.

**Properties.** It is a deep red, heavy liquid: it is the only element liquid at the ordinary temperature, except quicksilver. It has a strong smell, like chlorine, and is poisonous. It is slightly soluble in water, and the solution possesses bleaching powers. It boils at  $145.4^{\circ}$  F.; at  $9.5^{\circ}$  F. it becomes a red crystalline solid. It stains the skin yellow.

Like chlorine, it combines with hydrogen and forms **hydrobromic acid**, or hydric bromide ( $\text{HBr}$ ), which is very soluble in water.

There are also several oxides of bromine, of which **bromic acid** ( $\text{HBrO}_3$ ) is the best known.

Bromine combines with metals, and forms **bromates**. These, when heated, part with oxygen, and become **bromides**.

---

## IODINE.

*Symbol I. Combining weight 127. Sp. gr. of vapour 8.716. Sp. gr. of solid 4.947.*

Iodine is found in sea-water, but in smaller quantities than bromine. It was discovered by Courtois in 1811. It has also been found combined with silver in Mexico, and zinc in Silesia.

**Preparation.** It is obtained from **kelp**, the ash of certain sea-weeds. The kelp is digested in water, and the liquor reduced by evaporation, until only the salts of iodine are left in solution. The iodine liquor is heated with black oxide of manganese and sulphuric acid, and iodine distils over as a deep violet-coloured vapour, which condenses to a bluish-black solid, with bright metallic lustre.

In imitation of the iodine liquor, a solution of iodide of potassium may be taken for the preparation of iodine.

**Properties.** It melts at  $225^{\circ}$  F. and boils at  $347^{\circ}$  F. It smells something like chlorine; it stains the skin a brown

colour. Taken internally it is poisonous; but it is used in medicine for glandular swellings. It is slightly soluble in water, to which it gives a yellow colour. It is called  $\text{Ιωδης}$  (Iodes)—violet-coloured, from the colour of its vapour.

It combines with many of the metals, forming **iodides**, which are decomposed by chlorine.

Iodine forms with starch a beautiful blue compound.

**HYDRIODIC ACID (HI)** is a compound of iodine and hydrogen, and is obtained by heating these together. It may also be obtained more conveniently by applying heat to a glass tube, containing alternate layers of iodine, powdered glass, and chips of phosphorus.

A solution of hydriodic acid is obtained by sending sulphuretted hydrogen through water containing iodine.

It dissolves many of the metals, giving off hydrogen, and forming metallic iodides.

#### TESTS FOR IODINE.

1. Corrosive sublimate (bichloride of mercury,  $\text{HgCl}_2$ ) produces a bright scarlet precipitate of biniodide of mercury ( $\text{HgI}_2$ ), which is soluble in excess.

2. Nitrate of silver ( $\text{AgNO}_3$ ) gives a pale primrose yellow precipitate of iodide of silver, insoluble in nitric acid.

3. Acetate of lead gives a bright yellow precipitate of iodide of lead ( $\text{PbI}_2$ ), soluble in much cold water.

4. The solution, heated with binoxide of manganese and sulphuric acid, gives off iodine, which may be known from its violet-coloured vapour.

#### COMPOUNDS OF IODINE AND OXYGEN.

Iodine has a greater attraction for oxygen than chlorine or bromine, and forms with it two oxides, viz, **iodic acid** ( $\text{HIO}_3$ ) and **periodic acid** ( $\text{HIO}_4$ ).

**Iodic acid** may be prepared by boiling iodine in nitric acid. It combines with metals, and forms salts called **iodates**. It is useful as a test for morphia.

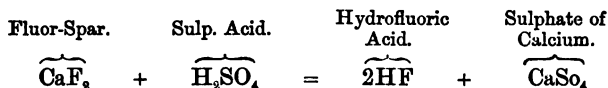
**Periodic acid** is a substance of little interest.

## FLUORINE.

*Symbol F. Combining weight 19.*

This element is not found free. It is found combined with the metal calcium, as fluoride of calcium ( $\text{CaF}_2$ ), commonly called **fluor-spar**. It forms no compounds with oxygen. Its most important compound is—

**Hydrofluoric acid.** This answers to the hydrogen compounds of chlorine, bromine, and iodine, and may be prepared in a similar manner. Sulphuric acid is poured upon fluor-spar in the state of fine powder; heat is applied, and hydrofluoric acid comes off as dense fumes.



Hydrofluoric acid attacks glass, therefore it must be prepared in metallic vessels. Its fumes are very corrosive. The smallest drop of the liquid acid upon the skin produces a deep and painful wound. Its most interesting property is that of etching glass, on account of the affinity of fluorine for the silicon contained in glass. This property may be shown by coating a piece of glass with wax, and, after removing a portion of the wax with a sharp point, exposing it to the fumes of the acid.

When this acid is poured into water, it is dissolved, making a hissing noise.

Fluorine forms an interesting compound with silicon, called the **fluoride of silicon** ( $\text{SiF}_4$ ). It is formed whenever free hydrofluoric acid comes in contact with silicon. It is a colourless gas, decomposed by water. Each bubble of gas, as soon as it enters water, is changed into a bladder of gelatinous silver. The gas may be prepared by heating in a flask equal parts by weight of finely powdered fluor-spar and white sand, and eight parts of sulphuric acid, and collecting it over mercury.

#### QUESTIONS ON IODINE, BROMINE, FLUORINE.

1. Describe Iodine. (O.)
2. What is the principle source from which Iodine is obtained? Describe and explain the process for extracting it, supposing it to be in the state of Iodide of Sodium. (C.)
3. Whence is Bromine obtained, and by what process? (O.)
4. State the composition and the mode of preparing Hydrofluoric Acid. \*
5. The Iodine in 100 volumes of Hydriodic Acid is liberated in succession by Chlorine and by Oxygen. How many volumes of Chlorine, and how many volumes of Oxygen are required? \*
6. One cubic foot of Hydriodic Acid is decomposed by an excess of Bromine, how many cubic feet of Hydrobromic Acid are formed? \*

---

#### SULPHUR.

*Symbol S. Combining weight 32. Sp. gr. of vapour at 900°, 6.617.*

Sulphur is found in nature both free and combined. It occurs free in beds in Sicily, Iceland, and volcanic districts generally. It is found combined with many metals as sulphides: thus—iron, copper, lead, zinc, are commonly found in combination with it. It is also found in combination with various earths in the state of sulphuric acid; these compounds are called sulphates, as the sulphate of calcium, magnesium, &c.

Sulphur is obtained pure by heating a sulphide, as sulphide

of iron (iron pyrites), in earthenware pots; the sulphur distils over in the form of vapour, and is condensed.

It occurs in commerce as roll sulphur, or brimstone, and flowers of sulphur; the latter is the vapour of sulphur quickly condensed into a very fine powder.

**Properties.** Sulphur is highly inflammable; it burns in the air with a blue flame, combining with oxygen to form sulphurous acid, well known from its suffocating odour. It is insoluble in water. It is a bad conductor of heat,—hence, when held in the hand, it crackles and splits.

It melts at 239° F., and forms a yellow liquid, which becomes darker as the heat is increased. At about 450° F. it is nearly solid; but at a still higher temperature it becomes liquid again.

Sulphur is largely used in the arts. It is one of the components of gunpowder. It is also used in medicine. In the form of sulphurous acid it is used in bleaching silks and flannels, &c.

#### COMPOUNDS OF SULPHUR AND OXYGEN.

Sulphur forms seven compounds with oxygen, but only three of these are of importance, viz :—

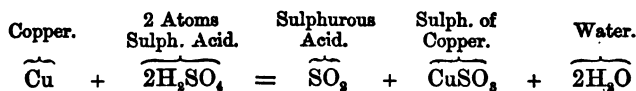
Sulphurous Acid	...	...	SO <sub>2</sub>
Sulphuric Acid	...	...	SO <sub>3</sub>
Hyposulphurous Acid	...	...	S <sub>2</sub> O <sub>3</sub>

The other four compounds are, as yet, little known.

**SULPHUROUS ACID**, or Sulphurous Anhydride. *Symbol* SO<sub>2</sub>. *Combining weight* 64. *Sp. gr.* 2.247.

When sulphur burns, this oxide is produced, which is known from its peculiar smell. It issues in large quantities from the craters of volcanoes. When it is required pure, it is obtained by depriving sulphuric acid of one atom of oxygen. This is

done on the small scale by heating copper with strong sulphuric acid. The latter gives one-third of its oxygen to the copper, and becomes sulphurous acid, which passes off as a gas. The undecomposed sulphuric acid remains behind, in union with the copper, as copper sulphate. Thus :—



**Properties.** It is a colourless, invisible gas, very soluble in water. Water dissolves at 32° F. 68·8 times its volume of the gas : at 59° F. 43·5 times its bulk. Such a solution is called **hydrated sulphurous acid**, and has the symbol  $\text{H}_2\text{SO}_3$ . On account of its great weight, the gas may be collected by displacement of air. It is incombustible, and a non-supporter of combustion.

It possesses bleaching powers, and is used in bleaching straw, wool, silk, isinglass, sponge. It has also disinfectant properties. Sulphurous acid forms with metals, salts, called **sulphites**.

**SULPHURIC ACID**, or Sulphuric Anhydride. *Symbol*  $\text{SO}_3$ .  
*Combining weight* 80.

This, the most important of all the acids, does not occur in nature free, but it is found in those compounds called sulphates. The ordinary acid of commerce is a compound of sulphuric anhydride and water, and its symbol is  $\text{H}_2\text{SO}_4$ . It is made in England to a large extent, especially in South Lancashire.

**Preparation.** Ordinary sulphuric acid is obtained by adding to sulphurous acid one atom of oxygen, by means of nitric acid and steam. This process is as follows :—Sulphur (or iron pyrites) is burnt in a furnace, and forms sulphurous acid. This is conveyed into large leaden chambers, along with nitric acid and steam. The nitric acid parts with a portion of its oxygen to the sulphurous acid, and the sulphuric acid thus formed

condenses on the floor of the chamber in combination with water, and is drawn off.

The nitric acid, on parting with its oxygen, is changed to nitric oxide, and this, as soon as formed, combines with the oxygen of the air freely admitted into the chamber, and becomes hyponitric acid. This hyponitric acid again parts with its oxygen to the sulphurous acid, and returns to the state of nitric oxide. Thus these compounds of nitrogen act as carriers of oxygen to sulphurous acid.

As the sulphuric acid, when first drawn off, is too weak, it is evaporated in leaden pans until it has the sp. gr. of 1.848. It then forms the oil of vitriol of commerce.

In Nordhausen, in Saxony, sulphuric acid is prepared by heating the sulphate of iron (green vitriol). This acid goes by the name of Nordhausen sulphuric acid, and when strongest, it is twice the strength of the English acid.

SULPHURIC ANHYDRIDE ( $\text{SO}_3$ ) is obtained from the **Nordhausen** acid by gently heating the latter, when white fumes are given off, which condense into soft fibres, tough and ductile, and can be moulded in the fingers like wax. This substance has no acid properties, but it dissolves in water with a hissing noise like a red-hot iron, and gives to the solution the properties of ordinary sulphuric acid.

**Properties.** The sulphuric acid of commerce, commonly called **oil of vitriol**, is a dense, colourless, oily fluid. It has a great attraction for water, and when mixing with it, gives off great heat. This property of the acid is made use of in drying substances that cannot be exposed to heat. It has a strong sour taste, reddens vegetable-blues, and changes browns to yellows. It is corrosive, destroying all organic substances it touches.

It dissolves most of the metals when heated. Zinc, iron, cobalt, nickel, manganese, are dissolved by dilute sulphuric acid, when hydrogen is given off, and sulphates of the metals



are formed. It takes the place of every other acid in a compound which boils at a lower temperature than itself. For example, nitric acid boils at  $184^{\circ}$  F., sulphuric acid boils at  $640^{\circ}$  F.,—therefore, when the latter acid is heated with a nitrate, it drives off the nitric acid, and takes its place.

#### TESTS OF SULPHURIC ACID.

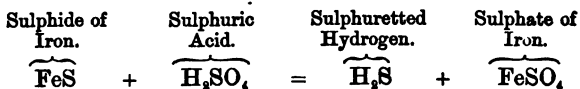
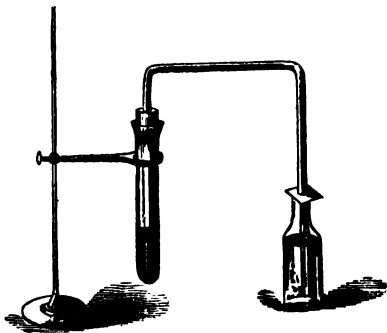
Sulphuric acid and all its salts are recognised by giving with chloride of barium a white precipitate, insoluble in nitric acid.

**HYPOSULPHUROUS ACID** ( $\text{H}_2\text{S}_2\text{H}_2\text{O}_4$ ) is an acid of importance, because it forms, with soda, a salt—**hyposulphite of soda** ( $\text{Na}_2\text{S}_2\text{H}_2\text{O}_4 + 4\text{H}_2\text{O}$ ), which is used in photography for fixing images. The photograph is washed in a solution of this salt, when the silver salt with which the glass was previously covered, is completely dissolved, whilst the part blackened by the light is left untouched.

#### COMPOUNDS OF SULPHUR AND HYDROGEN.

**HYDROSULPHURIC ACID**, Sulphuretted Hydrogen, or Hydrogen Sulphide, (*Symbol*  $\text{H}_2\text{S}$ —*Combining weight* 34—*Sp. gr.* 1.191), is an important compound of sulphur and hydrogen.

**Preparation.** It is generally prepared by adding diluted sulphuric acid to the sulphide of iron ( $\text{FeS}$ ). The iron and hydrogen change places; sulphuretted hydrogen comes off as a gas, and sulphate of iron or green vitriol remains. Thus:—



**Properties.** It is a colourless gas, having the smell of rotten eggs. It is very poisonous when respired; a small quantity is sufficient to destroy the lower animals. It burns with a pale, bluish flame, forming sulphurous acid and water. It is very soluble in water, which at 32° F. dissolves 4·37 volumes of the gas. The solution possesses the properties of the gas, but, after a while, it is decomposed, depositing its sulphur.

The gas is immediately decomposed by chlorine, bromine, and iodine, the hydrogen combining with these substances.

Sulphuretted hydrogen occurs in certain springs, as, for instance, in Harrogate waters.

Sulphuretted hydrogen is of constant use in the laboratory as a test for various metals. If this gas be added to acid solutions of arsenic, antimony, tin, gold, platinum, silver, mercury, lead, bismuth, copper, and cadmium, it immediately precipitates sulphides of these metals. If, however, it be added to acid solutions of iron, chromium, aluminium, cobalt, manganese, nickel, or zinc, no precipitates are formed, because the sulphides of these metals are soluble in acid solutions; but they are precipitated on the addition of ammonia:—thus this gas serves to divide the metals into groups. Most of the above precipitates are black. The precipitate in the arsenic solution is **yellow**; in antimony—**orange-coloured**; in zinc—**white**; in manganese—**flesh-coloured**.

As this gas is present in small quantities in coal-gas, the use of the latter is injurious to paintings, and tarnishes the silver goods of jewellers, &c.

#### TESTS FOR SULPHURETTED HYDROGEN.

1. Most of the sulphides or hydrosulphates are known by the smell of sulphuretted hydrogen, which they give off when moistened with hydrochloric acid.

2. A sulphide before the blow-pipe gives off the smell of sulphurous acid.

**CARBONIC DISULPHIDE**, or Bisulphide of Carbon. *Symbol* CS<sub>2</sub>. *Combining weight* 76. *Sp. gr.* 1.27.

This compound is formed by passing vapour of sulphur over red-hot carbon. It is an oil, insoluble in water. It has the smell of decaying vegetable matter, and is very poisonous. It is used as a solvent for many oils and fats. Gums, caoutchouc, sulphur, phosphorus, dissolve in it. It burns with a blue flame, producing sulphurous acid and carbonic acid.

#### QUESTIONS ON SULPHUR AND ITS COMPOUNDS.

1. Describe the chief allotropic modifications of Sulphur, and name the chief substances from which Sulphur is obtained. How is Sulphurous Acid obtained? How much Sulphur, and how much Oxygen, are contained in 100 grammes of Sulphurous Acid? (L.)
2. Describe the process of manufacturing Sulphuric Acid, and state its most important uses. (O.)
3. By what different methods can the oxidation of Sulphurous Acid be effected? Explain the principles on which they severally depend? (O.)
4. Describe the method of obtaining Sulphuretted Hydrogen. For what purposes is it employed? Mention some cases in which it is naturally produced. What means would you employ to remove it from a tainted atmosphere, and how does the substance you name act upon it chemically? (C.)
5. You have given to you iron filings, Sulphur, and Hydrochloric Acid, and are required to make Sulphuretted Hydrogen, how will you proceed? \*
6. What compound is produced when Sulphur is burnt? How may this be converted into Sulphuric Acid? (C.)
7. How much Oxygen, Sulphur, Hydrogen, and Nitrogen are contained in 100 grains of H<sub>4</sub>NOSO<sub>3</sub>? (C.)

---

#### SELENIUM.

*Symbol* Se. *Combining weight* 79. *Sp. gr. of solid* 4.78.

This is a rare element, discovered by Berzelius, in 1817, and

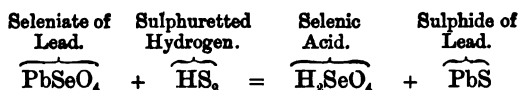
called by him Selenium, from the Greek word  $\sigma\epsilon\lambda\eta\nu\eta$  (Selene), the moon. It has properties much like sulphur, and is always found in combination with metals, as **selenides**.

It may be obtained in three states, viz., amorphous, vitreous, and crystalline. At ordinary temperatures it is a solid of a deep brown colour, and metallic lustre. It has neither taste nor smell, and is insoluble in water. When heated in closed vessels, it gives off a deep yellow vapour. It burns in air with a bright blue flame, giving off a smell of rotten cabbage.

With oxygen, selenium forms two compounds, which, combined with water, are acids, and correspond to the sulphur compounds, sulphurous and sulphuric acid.

**SELENIOUS ANHYDRIDE** ( $\text{SeO}_2$ ) is formed when selenium is burnt in air or oxygen. When combined with one atom of water, it becomes **selenious acid** ( $\text{H}_2\text{SeO}_3$ ).

**SELENIC ACID**, or Dihydric Seleniate ( $\text{H}_2\text{SeO}_4$ ), is best prepared by fusing selenium or a selenite with nitre, and adding nitrate of lead to a solution of the mass. An insoluble salt of seleniate of lead is formed, which may be dissolved by sulphuretted hydrogen, yielding dihydric seleniate and sulphide of lead. Thus :—



Selenic acid has properties resembling sulphuric acid, and its salts, called **seleniates**, correspond to the **sulphates**. It is the only acid which dissolves gold. On platinum it has no action.

With hydrogen, selenium forms a compound—seleniuretted hydrogen ( $\text{H}_2\text{Se}$ ), which may be obtained by the action of an acid upon a selenide. It has a most disagreeable smell, and if respired in the smallest quantity, it causes great irritation.

## SILICON.

*Symbol* Si. *Combining weight* 28. *Sp. gr.* 28.

Next to oxygen this is the most abundant element in nature. It is always found combined with oxygen, forming silica. This compound forms the bony skeleton of the earth. Rock-crystal, sand, flint, and many minerals are chiefly composed of silica; and, because of its abundance in flint or *silex*, it receives its name.

Silicon may be obtained in three states, viz :—

(1.) As a brown amorphous powder, by heating the silicofluoride of potassium with potassium. This burns brilliantly in air, and is changed into silica.

(2.) In the graphite form, like the graphite form of carbon, by heating the amorphous powder to a high temperature.

(3.) As a crystalline form, by fusing the mixture, which gives amorphous powder, with zinc. Crystals of silicon are formed on the zinc, and the latter may be removed by solution in an acid. The crystals are hard enough to cut glass.

SILICIC ACID ( $\text{SiO}_2$ ), or *Silica*, is the only known oxide of silicon. It occurs pure in rock-crystal, white sand, and some sorts of quartz. It is found less pure in sandstone, flint, felspar, and a variety of minerals. It enters largely into the composition of many precious stones, as agate, amethyst, onyx, opal.

Silica may be obtained as a powder, as follows :—One part of finely divided quartz or white sand is heated with four parts of carbonate of soda; the mixture fuses, and gives off carbonic acid; the remainder is boiled with water, and dissolves, when, on the addition of hydrochloric acid, hydric silicic acid ( $\text{H}_4\text{SiO}_4$ ) separates as a gelatinous mass. If this be evaporated to dryness, then heated, and hydrochloric acid added, silicic acid is formed as a white powder.

**Properties.** Silica is a white, tasteless powder, insoluble in water and acids; but it dissolves in strong alkaline solutions. Its sp. gr. in the crystallised form is 2.64, in the amorphous form 2.2. It can only be fused at the oxyhydrogen flame.

In the gelatinous form it is soluble in water, but the solution has no action upon vegetable colours. It combines with bases, and forms compounds called **silicates**. It enters largely into the composition of glass, porcelain, and earthenware.

Silicon forms compounds with hydrogen, sulphur, chlorine, and fluorine. Compounds of the latter have been described under fluorine; the others are of little importance.

---

## BORON.

*Symbol B. Combining weight 10.9. Sp. gr. 2.68.*

Boron resembles, in properties, carbon and silicon. It is a body sparingly found in nature, and it is always combined with oxygen as **boracic acid**. From this acid boron may be obtained in **three** forms,—crystalline, graphitoidal, and amorphous. The amorphous form is obtained by heating boracic acid with potassium or sodium; the crystalline—by heating the amorphous form with aluminium; the graphitoidal form is also obtained in the same way.

Boron, if heated in air or in oxygen, takes fire below redness, and forms boracic acid. It also burns freely in chlorine.

**BORACIC ANHYDRIDE**, or Boracic Acid ( $B_2O_3$ ). *Combining weight 69.8.*

This is the only oxide of boron. It is found free in the lagoons of Tuscany; and combined with sodium, as **borax** or **tincal**, it is found in Thibet and California. A borate of lime and magnesia comes from South America.

Boracic acid is prepared by adding sulphuric acid to a hot solution of borax, when sulphate of sodium is formed, and hydrate boracic acid separates in crystals. On heating, the crystals part with the water, and pass into a fused glassy mass.

Boracic acid is slightly soluble in water. Its taste is faintly sour; it turns turmeric paper brown, and gives to litmus a reddish tint. It dissolves in alcohol, and the solution burns with a green flame.

Borax, the baborate of sodium, is the most important compound of boracic acid. Its symbol is  $\text{Na}_2\text{B}_4\text{O}_7 + 10\text{H}_2\text{O}$ . This substance is useful in the manufacture of the finer kinds of glass, and in forming glazes for porcelain.

Fused borax has the property of dissolving many metallic oxides, and forming, during the process, coloured glasses. This property makes this substance very useful in the laboratory for detecting certain oxides. Thus, an oxide of chromium gives, with borax, a bead of green colour; an oxide of copper, one of red; an oxide of manganese, one of violet.

Boron forms a compound with chlorine, called a tri-chloride ( $\text{BCl}_3$ ), and a similar one with fluorine ( $\text{BF}_3$ ).

---

## PHOSPHORUS.

*Symbol P. Combining weight 31. Sp. gr. 1.83.*

Phosphorus, which was discovered by Brandt of Hamburg in 1669, is not found in nature free; but in the state of phosphoric acid it occurs in primitive and volcanic rocks. From the soil formed by these rocks, plants extract phosphorus, and these latter supply the animal system with this substance. The bones, the brain, the nerves, and various tissues of animals, have phosphorus as one of their most important ingredients.

**Preparation.** The ashes of bones, which chiefly contain

phosphate of calcium, are acted upon by two-thirds of their weight of diluted sulphuric acid. The acid dissolves a part of the calcium (lime), forming sulphate of calcium, which separates as an insoluble white powder. The remaining liquid contains in solution the phosphorus, combined with calcium and oxygen, as superphosphate of calcium. This liquid is evaporated till it becomes as thick as syrup, and then mixed with charcoal and heated in an iron pot, till all the moisture is driven off. The mixture is afterwards heated in an earthenware retort, the neck of which dips under water. The phosphoric acid ( $P_2O_5$ ) of the superphosphate of calcium is decomposed by the charcoal; the oxygen unites with the charcoal to form carbonic oxide, and the phosphorus distils over, and condenses under water in yellow drops.

**Properties.** Phosphorus is a soft, slightly yellow solid, much resembling wax. When exposed to the air it gives off white vapours: it melts at  $111.5^{\circ}$  F., and boils at  $550^{\circ}$  F. It is very inflammable, and requires to be kept under water. It receives its name phosphorus (from  $\phi\omega\varsigma$ —light, and  $\phi\acute{\epsilon}\rho\omega$ —I bear) from its property of giving off, in a darkened room, a pale greenish light. It dissolves in oils and bisulphide of carbon, and from these solutions it may be obtained in a crystallized form.

Phosphorus may be obtained in various forms, but the most interesting is red, or **amorphous** phosphorus. This is obtained by heating yellow phosphorus in a vessel containing carbonic acid, at a temperature between  $446^{\circ}$  F. and  $464^{\circ}$  F. The red kind differs very much in its properties from the yellow kind. It does not take fire so readily, and therefore needs not to be kept in water; it is not soluble in oils, &c., and its specific gravity is 2.14.



## COMPOUNDS OF PHOSPHORUS AND OXYGEN.

Phosphorus and oxygen form the four following compounds—

				Combining Weight.
Oxide of Phosphorus	...	...	P <sub>4</sub> O	... 140
Hypophosphorous Acid	...	...	P <sub>2</sub> O	... 78
Phosphorous Acid (Phosphorous Anhydride)			P <sub>2</sub> O <sub>3</sub>	... 110
Phosphoric Acid (Phosphoric Anhydride)			P <sub>2</sub> O <sub>5</sub>	... 142

OXIDE OF PHOSPHORUS is said to be formed when a stream of oxygen is directed against phosphorus, melted beneath hot water.

HYPOPHOSPHOROUS ACID has never been obtained anhydrous. When phosphorus is boiled with a solution of potass or baryta, water is decomposed, giving rise to phosphuretted hydrogen, phosphoric acid, and hypophosphorous acid; the first escapes as gas, whilst the last two remain in union with the potass or baryta. The hypophosphorous acid, in union with the baryta, remains in solution, and, by filtering, may be separated from the insoluble phosphate. Sulphuric acid, added to the solution, precipitates the baryta, and leaves the hypophosphorous acid in solution. By evaporation it may be reduced to the consistency of syrup.

It forms salts called hypophosphites, which are all soluble in water.

PHOSPHOROUS ACID is obtained anhydrous by burning phosphorus in a limited supply of air, when a white powder is formed. The hydrated acid ( $H_3PHO_3$ ) is obtained by adding water to the terchloride of phosphorus ( $PCl_3$ ), when mutual decomposition takes place; the oxygen of the water unites with the phosphorus, forming phosphorous acid, and its hydrogen, with the chlorine, forming hydrochloric acid. The latter

is expelled by heating the liquid, and the phosphorous acid is obtained as crystals. This acid exposed to the air absorbs oxygen, and becomes phosphoric acid.

Its salts are called **phosphites**, but they are of little importance.

**PHOSPHORIC ACID** is obtained anhydrous by burning phosphorus in oxygen, or dry atmospheric air, when a snow-white powder is formed, which rapidly absorbs moisture, and runs into a liquid. In the hydrate state its symbol is  $H_3PO_4$ . Once in this state, the water cannot be separated from it.

Phosphoric acid is a powerful acid; it has a strong, sour taste, reddens litmus paper, and forms an important class of salts called **phosphates**. It is not poisonous.

Most acids unite with bases to form salts, in the proportion of one equivalent of the acid and one of the base, so that they are called monobasic, or single based. But phosphoric acid differs from these in possessing **three** modifications or forms, each having its own peculiarities; and these form salts by uniting with **one**, with **three**, or with **four** equivalents of the base. These modifications of phosphoric acid are distinguished as follows:—

Metaphosphoric Acid	unites with 1 equivalent of base.
Ortho-phosphoric Acid	„ 3 „
Pyro-phosphoric Acid	„ 4 „

## PHOSPHURETTED HYDROGEN.

*Symbol*  $PH_3$ . *Combining weight* 34. *Sp. gr.* 1.185.

Phosphuretted hydrogen, as a gas, may be obtained by heating hydrated phosphorous acid; but it is generally prepared by heating phosphorus and sticks of potass in a little water, which

produce phosphuretted hydrogen and hypophosphite of potassium.

Hypophosphite of  
Potassium.



Great care is required in this process. The retort should be small, and have a long tube, which should be placed under water from the first, as the gas formed in this way has the property of taking fire spontaneously in atmospheric air. Each bubble of the gas, as it escapes into the air, bursts into flame, and changes into a beautiful white wreath of phosphoric anhydride, expanding and revolving as it rises.

The inflammability of the gas prepared in this way is due to the presence of a minute quantity of the vapour of the liquid phosphide of hydrogen ( $P_2H_4$ ). The gas obtained by heating hydrated phosphorous acid is not inflammable.

Phosphuretted hydrogen has a strong smell of garlic; it is highly poisonous, and slightly soluble in water. When sent through solutions of certain metallic salts, it precipitates phosphides of the metals. It has no acid properties like sulphuretted hydrogen. Phosphuretted hydrogen is also known by the names—"Will-o'-the-Wisp" and "Jack-o'-Lantern."

There are three compounds of phosphorus and hydrogen, viz.—the gas described above; liquid phosphide of hydrogen ( $P_2H_4$ ); and solid phosphide of hydrogen ( $P_3H$ ).

#### QUESTIONS ON PHOSPHORUS AND ITS COMPOUNDS.

1. How much Anhydrous Phosphoric Acid is formed by the combustion of one gramme of Phosphorus ( $P = 31$ )? How much pure air is needed for the production of this Phosphoric Acid? (L.)
2. What combinations does Phosphorus form (1) with Hydrogen, and (2) with Chlorine? How is each of these substances obtained? (O.)
3. What are the characteristic properties of Phosphorus? How is it obtained? (O.)
4. Give the composition of Phosphorous Acid, and describe the method of preparing it. State also the composition of the Phosphites. \*

5. Give the formula of Phosphoretted Hydrogen, its percentage, composition by weight and by volume, and its specific gravity. \*
6. Describe the preparation of red phosphorus, and give its properties. (C.)

## MISCELLANEOUS QUESTIONS.

1. What are the sources and combinations of the element Boron? (O.)
2. How has Silicon been obtained? (O.)
3. What is Borax? Explain its use in chemical analysis. (C.)

## THE METALLIC ELEMENTS.

---

The metallic elements are much more numerous than the non-metallic, and they are distinguished from the latter by three characteristics, which are only shown, however, when metals are in mass:—(1) their power of reflecting light, causing the glittering appearance, which is termed the **metallic lustre**; (2) their power of conducting heat well; and (3) the power of conducting electricity.

All the metals are solid at ordinary temperatures, excepting mercury, which is liquid. They differ from each other in physical and chemical properties, and in uses.

**Physical properties.** One striking property of the metals is their lustre.

In colour there is considerable uniformity. In most cases it is a shade of white. One metal, **copper**, is **red**; gold, calcium, barium, and strontium, are **yellow**.

Their **specific gravity** differs widely, passing from potassium and sodium, which are lighter than water, to platinum, which is twenty-one times heavier.

Their **fusibility**, or melting point, is also widely different. **Mercury** melts at 39° F. below zero, whilst that of **iron** is nearly 3000 times higher, and that of **platinum** not exactly known.

The property of **malleability**, or the power of being hammered out into thin sheets, is possessed by certain of the metals. The most malleable is **gold**, which may be beaten out to the thinness of  $\frac{1}{880000}$  of an inch.

**Ductility**, or the power of being drawn out into wire, is a property of the malleable metals, and in this there is much difference. Gold, silver, platinum, iron, copper, are all nearly equally ductile.

**Volatility**, or the power of being converted into vapour, is possessed by many of the metals. Mercury boils at  $662^{\circ}$  F. ( $350^{\circ}$  C.); arsenic passes into vapour before becoming liquid; potassium, sodium, zinc, magnesium, and cadmium rise in vapour at a red heat.

**Hardness, brittleness, and tenacity** are important and widely different properties of the metals.

**Crystallization.** Almost all substances capable of existing in the solid state, assume a definite geometrical form, or are said to **crystallize**. Bodies which do not crystallize are called **amorphous**. The most beautiful examples of crystallization are seen in the natural minerals. As a rule, every substance possesses a definite form in which it always crystallizes, and by which it may be known.

The most common forms are the cube (1), and the regular octohedron (2),—the latter is an eight-sided solid, of which the faces are equilateral triangles.

Fig. 1.

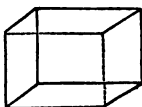
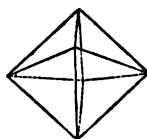


Fig. 2.



The many thousand different crystals are arranged into **six classes** or systems.

Certain substances exhibiting a similar chemical constitution, and crystallizing in the same forms, are called **isomorphous**. Substances which are found to crystallize in two different systems are called **dimorphous**.

## CHEMICAL PROPERTIES OF METALS.

Metals combine with one another to form alloys, and with the non-metallic elements, to form oxides, sulphides, chlorides, &c. In the latter case, the metallic characters are often lost, whilst in the former case they are preserved.

**ALLOYS.** As only few of the metals possess qualities which render them suitable to be employed alone, other metals are combined with them, and the compound is called an alloy. Thus gold and silver are hardened by an addition of a little copper. The hard and useful substance brass is an alloy of copper and zinc. Bronze is an alloy of copper and tin. Printing type is an alloy of lead and antimony.

When the metals combine with mercury, the resulting body is called an **amalgam**.

**METALLIC OXIDES.** Most metals readily combine with oxygen, and form oxides. The majority of metals burn in air, or pure oxygen; the precious metals, as silver, gold, and platinum, are not combustible, but combine with oxygen indirectly.

The oxides, as pointed out under oxygen, are divided into three classes, **basic**, **neutral**, and **acid**. Many oxides are found native. Others are obtained by combustion in air or oxygen, or by the decomposition of a salt.

**METALLIC SULPHIDES.** All the metals combine with sulphur to form sulphides, which are frequently found in nature as ores. Many metals unite with sulphur by heating them together; other sulphides are formed by heating the sulphate of the metal with charcoal; others are formed by the action of hydrosulphuric acid upon the metallic oxides.

The sulphides of the alkalis and alkaline earths are soluble in water; the remainder are insoluble, but some are soluble in acids and alkalis.

**METALLIC CHLORIDES.** Many metals combine directly with chlorine by bringing them together. Chlorides are found native, as common salt, (chloride of sodium). A very convenient way of obtaining a chloride is to add hydrochloric acid to a metallic oxide.

Metals combine with bromine, iodine, fluorine, &c., and these much resemble their corresponding chlorides.

Metals also unite with nitrogen, phosphorus, boron, carbon, silicon, and hydrogen,—but these compounds are not important.

#### DISTRIBUTION AND CLASSIFICATION OF METALS.

Only a few of the metals are found free in nature. Most of them are found in combination with oxygen, sulphur, or other metals. The compounds from which man extracts metals are called ores. These ores are usually found in the form of veins, which are cracks in the rocks composing the crust of the earth.

The metals may be conveniently grouped in four classes, viz. :—

- I. METALS OF THE ALKALIS.** (1) Potassium ; (2) Sodium ; (3) Lithium ; (4) Cæsium ; (5) Rubidium ; (6) Ammonium.

These are distinguished by forming with oxygen basic oxides, very soluble in water, and possessing alkaline properties.

- II. METALS OF THE ALKALINE EARTHS.** (1) Barium ; (2) Strontium ; (3) Calcium.

Their oxides are earthy powders, having alkaline properties, but less soluble than Class I.

- III. METALS OF THE EARTHS PROPER.** (1) Aluminum ; (2) Glucinum ; (3) Yttrium ; (4) Erbium ; (5) Terbium ; (6) Cerium ; (7) Lanthanum ; (8) Didymium.

Excepting aluminum, all these are very rare. Their oxides are earthy powders, insoluble in water.



IV. **METALS PROPER**, or the metals with which we are most familiar, thirty-four in number.

This class may be further divided into—**common** metals, as iron, lead, copper, &c., which readily combine with oxygen, and tarnish in air; and **noble** metals, as gold, which have only a feeble attraction for oxygen, and do not tarnish in air.

The metal magnesium is sometimes classified with the alkaline earths; but in several particulars it stands in close relation to zinc and cadmium, and for this reason it is reckoned as one of the metals proper.

---

I. **METALS OF THE ALKALIS.**

(1) Potassium; (2) Sodium; (3) Lithium; (4) Cæsium; (5) Rubidium; (6) Ammonium.

(1) **POTASSIUM (Kalium).**

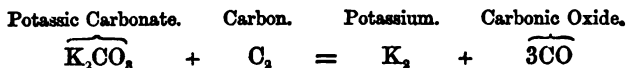
*Symbol K. Combining weight 39.1. Sp. gr. 0.865.*

This metal was discovered by Sir H. Davy, in 1807, by exposing potash to the action of a powerful galvanic battery.

Potassium is a bluish-white, bright metal. At ordinary temperatures it can be cut with a knife. At 32° F. it is brittle and crystalline; at 60°, soft; at 144.5°, it is liquid. It has great affinity for oxygen, so that if it is exposed to the air, it immediately becomes covered with a film of its oxide. On this account the metal must be preserved in some liquid, as naphtha, which does not contain oxygen. Its affinity for this gas is best seen by throwing a small piece into water, which it immediately decomposes. One atom of the metal liberates one-half the hydrogen of an atom of water, and forms the hydrate of potassium (caustic potass), KHO. The escaping hydrogen carries with it a small portion of the metal in the state of vapour, and

takes fire with the heat produced by the decomposition of the water, and burns with a beautiful purple flame, which is the colour of burning potassium.

**Preparation.** Potassium compounds occur in nature abundantly, especially in the igneous rocks, as granite, trap, &c. Plants extract these compounds from the soil formed from the crumbling of the rocks; and from the ashes of burnt plants, carbonate of potassium, (or potashes), is obtained. This carbonate is heated with charcoal, which removes the oxygen, produces carbonic oxide, and liberates the metal potassium. Thus—



Sea water is another source of potassium compounds. The nitrate and chloride of potassium are found in large quantities in various localities.

**OXIDES OF POTASSIUM.** Of these there are three—

1. Potash, or the Suboxide ... ..  $\text{K}_2\text{O}$
2. Deutoxide of Potassium ... ..  $\text{K}_2\text{O}_2$
3. Peroxide                   ,,       ... ..  $\text{K}_2\text{O}_4$

The two latter are of little importance, but the first, combined with water, is an important substance.

**POTASH** is procured pure by exposing thin slices of potassium to dry air. It is a white, fusible, and volatile solid, having great affinity for water. When once united with water, the latter cannot be separated by heat.

**POTASSIUM HYDRATE**, or Caustic Potash ( $\text{KHO}$ ), is obtained by boiling a solution of carbonate of potassium and quicklime: the lime combines with the carbonic acid, forming carbonate of lime, which falls as a heavy powder, and potassium hydrate remains in solution. The solution is evaporated to dryness, then the residue is melted and cast into moulds. It is sold in

the form of white sticks under the name of caustic potash. This is soluble in half its weight of water, and possesses strong caustic properties, destroying both animal and vegetable substances.

**SALTS OF POTASSIUM.** Potassium Carbonate, Potassic Carbonate, or Carbonate of Potassium ( $K_2CO_3$ ), known in commerce as potashes, or pearl-ashes, is brought in large quantities from America and Russia. The ashes of burnt wood are boiled with water, and the solution evaporated to dryness, when the residue forms the crude potashes; and these, when further purified, are called pearl-ashes. The pure carbonate may be obtained from these. A pure carbonate is also obtained by heating cream of tartar to redness, dissolving the mass in water, and evaporating the filtered liquid to dryness.

Carbonate of potassium is very soluble in water, and has strong alkaline properties. It is largely consumed in the arts, —in the manufacture of soap, glass, in preparing caustic potash, and other compounds of potash.

A bicarbonate of potassium ( $KHCO_3$ ) is formed by passing a current of carbonic acid through a solution of carbonate of potassium.

Potassium Nitrate, or Nitrate of Potassium ( $KNO_3$ ), well known as nitre, or saltpetre, is one of the most important salts of potassium. It is brought from the East Indies, where it is found in the soil. The soil is boiled with water, and the filtered solution evaporated, when the nitre crystalizes. It is also largely prepared artificially.

Saltpetre readily parts with its oxygen on the application of heat, and for this reason it is used in the manufacture of gunpowder.

Gunpowder is a mechanical mixture of nitre, sulphur, and charcoal, in the proportion of 1 atom of sulphur, 2 of nitre, and 3 of charcoal. When gunpowder is fired, the oxygen of the nitre unites with the charcoal, forming carbonic acid and car-

bonic oxide, the nitrogen is set free, and the sulphur unites with the potassium. The great explosive power of gunpowder is due to the sudden formation of these gases, which, on account of the heat produced, have a volume 1500 times greater than the volume of gunpowder employed.

Potassium forms several compounds with sulphur, of which the best known is the subsulphide ( $K_2S$ ).

The other important salts of potassium are **chloride of potassium** ( $KCl$ ),—a salt extracted from kelp, and largely used in the manufacture of alum: **iodide of potassium** ( $KI$ ),—a salt much used in medicine: **bromide of potassium** ( $KBr$ ): **chlorate of potassium** ( $KClO_3$ ): and **sulphate of potassium** ( $K_2SO_4$ ).

#### TESTS FOR POTASSIUM.

1. A solution of tartaric acid gives, after briskly stirring with a glass rod, a white precipitate of the acid tartrate of potassium (cream of tartar).

2. A solution of perchloride of platinum ( $PtCl_4$ ) produces a yellow crystalline precipitate of the chloride of platinum and potassium.

**Blow-pipe test.** The dry salt introduced upon platinum wire into the **reducing flame** of the blowpipe, communicates to it a violet tint. The colour of the flame may be seen by throwing saltpetre into an open fire.

#### (2) SODIUM (Natrium).

*Symbol* Na. *Combining weight* 23. *Sp. gr.* 0.972.

Sodium is found in nature abundantly, but always in a state of combination. Its most important compound is common salt, the chloride of sodium, which is found in Cheshire, Galicia, &c., and also in sea water. Compounds of sodium are found in several minerals, and in land and sea plants.

This metal was discovered by Sir H. Davy, about 1807. It has a silver-white colour, with a tinge of red. It is soft at

60° F., melts at 207·7° F.; it burns with a bright yellow flame. When thrown upon water, it decomposes a portion of it, like potassium, liberating the hydrogen, which, however, does not take fire unless the water be hot.

**Preparation.** Sodium is prepared from its carbonate in the same way as potassium. On account of its affinity for water, it must be kept in naphtha. It is much cheaper than potassium.

**OXIDES OF SODIUM.** There are two well-known oxides, viz.—

Oxide of Sodium, or Soda... $\text{Na}_2\text{O}$

Peroxide „ ...  $\text{Na}_2\text{O}_2$  (not used in the arts).

**Soda** is obtained anhydrous by burning sodium in dry air, a white powder being formed. It takes up moisture readily, and forms the hydrate ( $\text{NaHO}$ ).

The hydrate of sodium, or caustic soda, is obtained by boiling lime with the carbonate of sodium, and then evaporating the filtered solution to dryness. It is a white solid substance, highly soluble in water, caustic, alkaline, and closely resembles potash in its properties. It is largely used in soap-making.

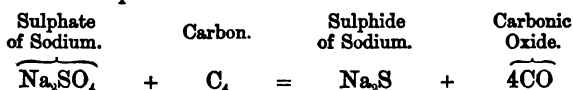
1. **SALTS OF SODIUM.** Chloride of Sodium, **common salt** ( $\text{NaCl}$ ), is the substance from which the sodium compounds are prepared. It is found in sea water, &c., as mentioned above. It is soluble in about three parts of cold water. It is largely used in the arts, and for domestic purposes.

2. **SULPHATE OF SODIUM**, or Glauber's salt ( $\text{Na}_2\text{SO}_4 + 10\text{H}_2\text{O}$ ), is obtained by the action of oil of vitriol upon common salt. It has a bitter taste, and is used in medicine.

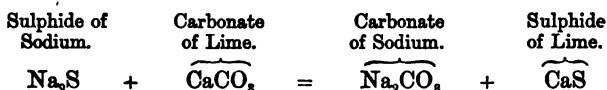
3. **CARBONATE OF SODIUM** ( $\text{Na}_2\text{CO}_3 + 10\text{H}_2\text{O}$ ), is a very important compound, largely used in glass-making, soap-making, bleaching, medicine, and for domestic purposes. It was formerly obtained from the ashes of sea plants, but it is now manufactured from the sulphate of sodium.

This sulphate, obtained from common salt by the action of

oil of vitriol, is heated in a furnace with chalk and powdered coal. Two stages of decomposition take place. In the first stage, the sulphate parts with all its oxygen to the coal, and becomes the sulphide: thus—



In the second stage, the sulphide parts with its sulphur to the lime of the chalk, whilst the carbonic acid of the chalk unites with the sodium of the sulphide, and becomes carbonate of sodium: thus—



The carbonate is separated by dissolving the mass in water, —then evaporating the carbonate solution to dryness. The residue is roasted, and is known as the **soda-ash** of commerce. If soda-ash be dissolved, and the solution allowed to stand, large crystals of the hydrated carbonate of sodium will be formed.

The carbonate is very soluble in water, and has strong alkaline properties.

**Bicarbonate of Sodium** ( $\text{NaHCO}_3$ ) is obtained by exposing the carbonate to the action of carbonic acid gas. A white powder is formed, of feeble alkaline action, used in medicine and for effervescing drinks.

**4. PHOSPHATES OF SODIUM.** There are several of these, but the one longest known, and from which most of the phosphates are formed, is the **hydro-disodic phosphate**, better known as rhombic phosphate of sodium ( $\text{Na}_2\text{HPO}_4 + 12\text{H}_2\text{O}$ ). It is a bitter, purgative salt, and used in medicine. It is also useful as a test.

**5. NITRATE OF SODIUM** ( $\text{NaNO}_3$ ), or **Cubic Nitre**, Chili saltpetre, is found in large beds in Peru. It is used in the manufacture of nitric and sulphuric acids, and also for manure.

Other compounds of sodium are **borax**—(see boron); the **sulphide** of sodium ( $\text{Na}_2\text{S}$ ); the **bromide** ( $\text{NaBr}$ ); the **iodide** ( $\text{NaI}$ ); and various **silicates**.

#### TESTS FOR SODIUM.

Nearly all the sodium salts are soluble in water, therefore there is no good direct test for this metal. In analysis, if the metal be proved to be one of the alkalis, the presence of sodium is shown by the absence of potassium and ammonium.

Before the blow-pipe, salts of sodium give a bright yellow colour to the flame.

#### (3) LITHIUM.

*Symbol* Li. *Combining weight* 7. *Sp. gr.* 0.59.

Lithium is of recent discovery, and is named from "lithos" ( $\lambda\iota\theta\omicron\varsigma$ )—a stone, because it was found at first only in minerals. It is a white metal, like sodium, and is the lightest solid known. It is widely distributed in nature; in small quantities it is found in micas and feldspars, in tobacco, and in mineral waters. The metal is obtained by decomposing its chloride by electricity.

Its oxide is called lithia ( $\text{Li}_2\text{O}$ ). This metal forms several compounds, as the chloride, the sulphate, the phosphate, and carbonate, all of which communicate a purplish red colour to flame.

#### (4) CÆSIUM.

*Symbol* Ca. *Combining weight* 133.

#### (5) RUBIDIUM.

*Symbol* Rb. *Combining weight* 85.36. *Sp. gr.* 1.52.

These metals were discovered about 1860; they are widely distributed, and found in small quantities in various rocks, mineral springs, and plants. They closely resemble potassium in their properties.

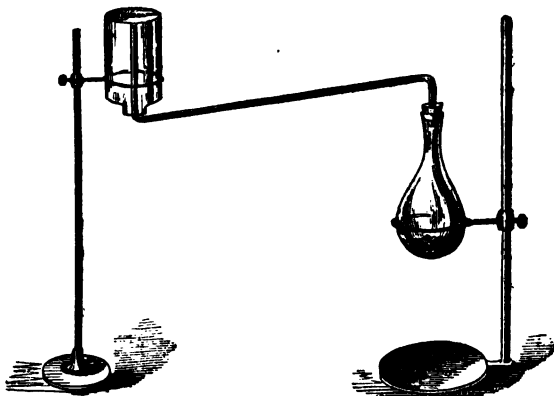
## (6) AMMONIUM.

*Symbol*  $\text{NH}_4$ . *Combining weight* 18.

As the salts of ammonium correspond in every respect to those of potassium and sodium, ammonium is classed under the head of the alkalis. The metal has never been isolated; but the body ( $\text{NH}_4$ ) forms an amalgam with mercury, and therefore it acts like a metal. Its oxide (or the compound resembling the oxides of the alkalis), ammonia, or hartshorn ( $\text{NH}_3$ ), is well known.

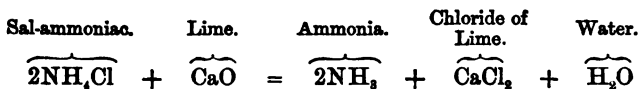
Ammonia is so called because it is readily obtained from a salt first found in Libya, near the temple of Jupiter Ammon, and hence termed **sal-ammoniac** (chloride of ammonium). It is called **hartshorn**, because the alchemists obtained it by distilling deer's horns. It is also called the volatile alkali.

**Preparation.** Most of the ammonia used in this country is obtained from the ammoniacal liquor produced in the manufacture of coal-gas. Hydrochloric acid is added to the liquor, and chloride of ammonium is formed. After evaporating to dryness, and further purification, the salt is ready for yielding ammonia.





Equal weights of quicklime and powdered chloride of ammonium are heated together in a retort, when ammonia comes off as a colourless gas, which may be collected by the displacement of air, or over mercury. One ounce of the salt yields about 750 cubic inches, or 30 grammes give about  $12\frac{1}{2}$  litres of the gas. The change may be shown thus:—



**Properties.** Ammonia has a strong, irritating smell, a bitter taste, and is feebly combustible. It is very soluble in water, and half as heavy as air, its sp. gr. being 0.59. If collected by the displacement of air, the vessel must be inverted, and the tube conveying the gas inserted till it almost touches the bottom of the vessel. It neutralizes the most powerful acids, and forms an important class of salts.

If any volatile acid be brought into an atmosphere containing ammonia, a white cloud is produced, on account of the union of the acid and ammonia to form a solid salt. A glass tube, moistened with an acid, will in this way detect the presence of ammonia.

Water at 32° F. absorbs 1050 times its volume of ammonia, at 59° F. 727 times its volume. The solution has all the properties of the gas; it is colourless, and strongly alkaline. All the ammonia may be driven off by heat.

### COMPOUNDS OF AMMONIUM.

The most important are as follows—

1. **SULPHIDE OF AMMONIUM** ( $\text{NH}_4\text{HS}$ ) is formed by passing sulphuretted hydrogen through a solution of ammonia to saturation. It is most useful in the laboratory as a test for the metals. When newly prepared it is colourless; it absorbs oxygen from the air, and becomes yellow. It has an offensive odour.

2. CHLORIDE OF AMMONIUM ( $\text{NH}_4\text{Cl}$ ) is the most important salt of ammonium. It is commonly called *muriate of ammonia* and *sal-ammoniac*. It is formed directly by the union of hydrochloric acid and ammonia gas; but it is obtained chiefly from the ammoniacal liquor of gas-works. It is usually sold in the shape of large hollow cakes, but it may also be obtained in crystals. It is very tough and fibrous, and very soluble in water. Its sp. gr. is 1.578.

3. SULPHATE OF AMMONIUM ( $2\text{NH}_4\text{SO}_4$ ) is obtained from ammoniacal liquor in the same way as the chloride, by adding oil of vitriol. It is used as a fertilizer; and it has lately been applied to muslins, &c., to prevent them burning with flame in case of contact with fire.

4. NITRATE OF AMMONIUM ( $\text{NH}_4\text{NO}_3$ ) is obtained by neutralizing nitric acid and ammonia. It is important as the source of nitrous oxide.

5. CARBONATES OF AMMONIUM. There are several compounds of carbonic acid and ammonium. The most important of these is the *sesqui-carbonate*, the common carbonate, or *smelling salts* of the shops,  $2(2\text{NH}_4\text{CO}_3)\text{CO}_2$ .

This carbonate is prepared by heating powdered chalk and chloride of ammonium. It has a strong pungent smell, like that of ammonia, an acrid taste, and a strongly alkaline reaction. It is much used in the arts and in medicine:—all the salts of ammonium may be obtained from it by adding the respective acids.

#### TESTS FOR AMMONIUM.

1. Ammonium is easily detected by its odour, which is given off from all its salts on the addition of quicklime or potash.

2. A glass rod dipped in hydrochloric acid produces dense white fumes when brought into the vapour of ammonia.

There is no blow-pipe test, as most of the salts of ammonium, when heated, are volatilized.

THE ALKALIS are distinguished from each other, thus:—

Tartaric Acid gives, with Potassium, a white precipitate.

„ „ Sodium, no precipitate.

„ „ Ammonium, no precipitate.

Ammonium and sodium are distinguished by the smell of ammonia on the addition of caustic potash.

The alkalis are distinguished from the alkaline earths by the following test:—Phosphate of sodium and ammonia produces in the latter white precipitates of the phosphates; in the former, no precipitate.

## II. METALS OF THE ALKALINE EARTHS.

There are three of these, viz.:—Barium, Strontium, and Calcium.

As Magnesium resembles these in some of its properties, we shall consider it under this head. These metals differ from the alkalis in forming insoluble carbonates and phosphates.

### (1) BARIUM.

*Symbol Ba. Combining weight 137.*

This metal was discovered by Sir H. Davy. in 1808. It has a yellowish-white colour, and occurs abundantly in the form of sulphate and carbonate.

The name Barium is derived from the Greek word βαρύς (heavy), on account of the great weight of its compounds.

OXIDES OF BARIUM. Of these there are two:—1. The protoxide baryta ( $\text{BaO}$ ). 2. The dioxide ( $\text{BaO}_2$ ). The last is of little importance.

**Baryta** ( $\text{BaO}$ ). This oxide is obtained anhydrous by heating the nitrate of barium in a crucible till the acid is expelled. It is a grey porous mass, slakes like lime when moistened with water, and forms the hydrate of baryta ( $\text{BaOH}_2\text{O}$ ). If hot water be saturated with this hydrate, and allowed to cool, the crystalline hydrate is formed ( $\text{BaOH}_2\text{O} + 8\text{H}_2\text{O}$ ). This hydrate is soluble in 20 parts of cold water:—the solution is alkaline and poisonous: and on exposure to the air absorbs carbonic acid and turns milky.

**SALTS OF BARIUM.** These are chiefly interesting as tests for the mineral acids, especially for sulphuric acid. The two most important are **chloride of barium** ( $\text{BaCl}_2$ ), and the **nitrate of barium** ( $\text{Ba}_2\text{NO}_3$ ). These are obtained by dissolving the carbonate or sulphide of barium in hydrochloric acid and nitric acid respectively. Chloride of barium is soluble in water, and is the usual test for detecting sulphuric acid in compounds. It detects the sulphates by forming white precipitates insoluble in nitric acid.

The other more important salts are the **sulphate of barium** ( $\text{BaSO}_4$ ), which is found native as **heavy spar**, a most insoluble salt, used as a paint; the **sulphide** ( $\text{BaS}$ ), formed from the sulphate by heating **heavy spar** with coal, and dissolving in water; the **carbonate** ( $\text{BaCO}_3$ ), which occurs native as the mineral **witherite**.

The carbonate and the soluble salts of barium are poisonous. The tests for barium will be considered with those of the other alkaline earths.

## (2) STRONTIUM.

*Symbol* Sr. *Combining weight* 87.5. *Sp. gr.* 2.54.

This metal resembles Barium in its properties. It was discovered about the year 1787, and was so called from Strontian, in Argyleshire, where it is found in the lead mines as a

carbonate and sulphate. It is less abundant than barium. It is a malleable metal, of a pale yellow colour, and is prepared by the action of electricity upon the fused chloride. It burns in air with a crimson flame, and forms the oxide strontia ( $\text{SrO}$ ).

Strontia may be obtained by heating the nitrate. It forms with water a hydrate ( $\text{SrO} + 9\text{H}_2\text{O}$ ), which is soluble in water, and has all the properties of baryta water.

**SALTS OF STRONTIUM.** The two most important are the chloride of strontium ( $\text{SrCl}_2$ ), and the nitrate of strontium ( $\text{Sr}_2\text{NO}_6$ ), which are soluble in water. They are used in the manufacture of fireworks to give a crimson colour to flame.

The salts found native are the carbonate ( $\text{SrCO}_3$ ), and the sulphate ( $\text{SrSO}_4$ ), known in mineralogy by the names of **strontianite** and **celestine** respectively. These native salts are insoluble in water.

### (3) CALCIUM.

*Symbol* Ca. *Combining weight* 40. *Sp. gr.* 1.58.

This metal is one of the most abundant and important constituents of the earth's crust. It is found in fluor-spar, in combination with fluorine; in limestone and chalk, in combination with carbonic acid; in gypsum, combined with sulphuric acid. The metal was discovered by Davy, in 1808; it is obtained from the chloride by electricity. It has a light yellowish colour, very malleable, and combines with oxygen on exposure to the air, forming its only oxide, **lime** ( $\text{CaO}$ ). The metal burns in air with a bright light. It derives its name **Calcium** from the Latin, **Calx** (lime).

**OXIDE OF CALCIUM, or Lime**, is obtained by heating limestone (the carbonate of calcium) in kilns; the carbonic acid is driven off, and lime remains. Pure lime, or quicklime, is a white, porous, infusible substance; it readily combines with water, giving off great heat, and forming the hydrate, or slaked lime

( $\text{CaOH}_2\text{O}$ ). The hydrate is soluble in about 700 parts of cold water, and in about 1300 of hot water, so that a solution of lime-water loses nearly half its lime on boiling. Lime-water has all the properties of solutions of potash and soda. It is used in the laboratory as a test for carbonic acid.

Lime is chiefly used in making mortar and cements. Mortar is a mixture of slaked lime and sand; lime, as it dries, absorbs carbonic acid from the air, and hardens on account of the formation of carbonate of calcium. Sand is used in the making of mortar, to prevent the lime crumbling and falling to pieces as it dries; but, besides this, a gradual combination takes place between the lime and silica of the sand, which contributes much to the hardness of mortar.

Ordinary mortar cannot be used under water. Mortars for this purpose, called hydraulic mortars, are obtained by using lime containing clay and silica.

Lime is also used as a manure. On rich vegetable soils it is useful in destroying the excess of vegetable matter; on heavy clay soils it liberates the alkalis, especially potash.

### COMPOUNDS OF CALCIUM.

CARBONATE OF CALCIUM ( $\text{CaCO}_3$ ) is found most abundant in nature, as chalk, limestone, marble, coral. It is the chief constituent of the shells of fish, egg shells, and coral reefs. It is found in the crystalline form as Calc spar, or Iceland spar. It is soluble in pure water to the extent of two grains in one gallon, but it is freely taken up by water containing carbonic acid. Water thus charged with carbonate of calcium is called **hard**. As the excess of carbonic acid escapes, or is expelled by heat, carbonate of lime is deposited. In this way, those masses of deposits called **stalactites**, found in limestone caves, are formed.

Most spring waters contain carbonate of lime, held in solution by excess of carbonic acid. During boiling, this acid is

expelled, and the well-known crust of carbonate of lime is formed inside the boiler. This evil is prevented by adding sal-ammoniac to the water, which produces the soluble chloride of calcium, and volatile carbonate of ammonium.

Hard waters may be softened by boiling, or by the addition of lime water to neutralize the excess of carbonic acid.

Permanent hardness of water is due to the presence of sulphate of calcium.

SULPHATE OF CALCIUM ( $\text{CaSO}_4$ ) is found **anhydrous** in a crystalline form in the Tyrol and Austria; but it is most abundant as a hydrate in combination with two atoms of water. In this state it is called **gypsum** and **alabaster**, and by the mineralogist **selenite** ( $\text{CaSO}_4 + 2\text{H}_2\text{O}$ ). When **gypsum** is carefully heated, it parts with its water, and forms **plaster of Paris**. This, when moistened, takes up again two atoms of water, and returns to a solid mass. Plaster of Paris is also called **stucco**, and is used for moulds and casts.

Sulphate of calcium is soluble in water to the extent of 150 grs. per gallon. The solution is useful for a test to distinguish the alkaline earths from each other. This salt is a common impurity of spring water, and gives to it the property of permanent hardness.

CHLORIDE OF CALCIUM ( $\text{CaCl}_2$ ) is obtained by dissolving limestone or marble in hydrochloric acid. By evaporation of the solution, the chloride is obtained in crystals associated with six atoms of water: four of these atoms are expelled by gentle heat. The salt, in this form, is used for freezing mixtures, and drying gases.

Chloride of calcium must not be confounded with **chloride of lime** ( $\text{CaOCl}_2$ ). The latter is known as **bleaching powder**, and is obtained by exposing the hydrate of lime, very slightly moist, to chlorine gas. The powder is also used as a disinfectant.

PHOSPHATES OF CALCIUM. Of these there are various com-

pounds, but the most interesting is **bone phosphate** ( $\text{Ca}_2\text{PO}_4$ ), which is the chief constituent of the bones of animals. This phosphate may be obtained by digesting calcined bones in nitric acid, and precipitating the filtered solution by ammonia.

**FLUORIDE OF CALCIUM, FLUOR-SPAR** ( $\text{CaF}_2$ ) is mentioned under fluorine. Other compounds are the **phosphide** ( $\text{CaP}_2$ ), and the **sulphide** ( $\text{CaS}$ ).

#### (4) MAGNESIUM.

*Symbol* Mg. *Combining weight* 24. *Sp. gr.* 1.74.

This metal is found abundantly in nature in combination with carbonate of lime, as magnesian limestone. It occurs in sea water as the chloride, and in many springs as the sulphate.

It is prepared by heating the **chloride** along with **sodium**, when chloride of sodium is formed, and metallic magnesium.

In properties, this metal closely resembles zinc and cadmium.

It is of a silver-white colour, malleable, and ductile. It does not tarnish in dry air; but it slowly oxidizes in moist air.

It burns in air with a dazzling white light, and, on account of this, it is used in photography as a substitute for sunlight.

**Magnesia** ( $\text{MgO}$ ) is the only known oxide of this metal. It is formed by burning magnesium in air, but it is usually obtained by heating the carbonate, or nitrate. It is a white, tasteless, soft, almost insoluble powder, with feeble alkaline properties. It forms a hydrate with water.

#### COMPOUNDS OF MAGNESIUM.

**SULPHATE OF MAGNESIUM** ( $\text{MgSO}_4 + 7\text{H}_2\text{O}$ ) is the most important compound. It may be obtained from sea water by evaporating the **bittern** along with sulphuric acid, or by the action of sulphuric acid upon magnesian limestone. This salt is a common ingredient in mineral waters, and because of its occurrence in the springs about Epsom, it is known as **Epsom**



**salts.** It is soluble in three times its weight of water ; by its solubility it is distinguished from the sulphates of the three preceding metals.

**CARBONATE OF MAGNESIUM** ( $\text{MgCO}_3$ ) is found native as a white, hard mineral, called **magnesite**. It is not easily prepared artificially.

The common white magnesia of the shops is a mixture of hydrated magnesia and the hydrated carbonate.

**CHLORIDE OF MAGNESIUM** ( $\text{MgCl}_2$ ) is interesting as being the compound from which the metal is obtained.

**PHOSPHATES OF MAGNESIUM.** The most important phosphate is that known as the **phosphate of magnesium and ammonium** ( $\text{MgNH}_4\text{PO}_4 + 6\text{H}_2\text{O}$ ). It is formed by adding any soluble salt of phosphoric acid, and ammonia, to a solution of sulphate of magnesium. This compound is soluble in pure water. The formation of this salt furnishes a delicate test for the presence of magnesium.

#### TESTS FOR THE ALKALINE EARTHS.

**Sulphate of Lime** produces, in a salt of **Barium**, a white precip. immediately.

„ „ **Strontium** a white precip. after a time.

„ „ **Calcium or Magnesium**, no precip.

**Oxalate of Ammonium** „ **Calcium**, a white precip.

„ „ **Magnesium**, no precip.

**With the blow-pipe**, **Barium** gives to a flame a **yellowish-green** colour.

„ **Strontium** „ **crimson** colour.

„ **Calcium** „ **orange red** „

**Magnesium** moistened with nitrate of cobalt on charcoal takes a pale flesh colour.

## RULES FOR SUNDRY ARITHMETICAL EXERCISES.

1. The Rule for correcting gases for Temperature and Pressure has been given on page 11.

2. The method of finding the volume of a substance produced in certain chemical decompositions, &c., is shewn on page 18.

3. To find the percentage composition and quantity of a composite body when its symbol or formula is given :—

- (i.) The percentage composition is found by multiplying the atomic weight of each element by the number of its atoms, this product again by 100, and dividing by the atomic weight of the given compound.
- (ii.) The quantity by weight in a given weight of a compound is found by multiplying the atomic weight of each element by the number of its atoms, and this product by the given weight of the compound.

Ex. 1. Find the percentage composition of Nitric Acid ( $\text{HNO}_3$ ).

Molecular weight of  $\text{HNO}_3 = (1 \times 1) + (14 \times 1) + (16 \times 3) = 63$ .

Since 63 parts of Nitric Acid contain 1 of Hydrogen, 14 of Nitrogen, 48 of Oxygen.

∴ 1 part of	"	"	$\frac{1}{63}$	"	$\frac{14}{63}$	"	$\frac{48}{63}$	"
∴ 100 parts of	"	"	$\frac{1}{63} \times 100 = 1.429$	"	per cent. of H	"	$\frac{48}{63} \times 100 = 76.19$	"
						"	N	
						"	O	

Ex. 2. How many grammes of Oxygen are contained in a kilogramme of Chlorate of Potassium ( $\text{KClO}_3$ )?

Molecular weight of Salt =  $\overset{\text{K}}{39.1} + \overset{\text{Cl}}{35.5} + \overset{\text{O}_3}{48} = 122.6$ .

122.6 grammes of  $\text{KClO}_3$  contain 48 grammes of Oxygen.

∴ 1 gramme	"	"	$\frac{48}{122.6}$	"	"
∴ 1 kilogramme,	}	"	$\frac{48}{122.6} \times 1000 = 391.52$	grms.	
or 1000 grammes					

4. To find the symbol or formula of a body, its percentage composition being given :—

Divide the given percentage number of each element by its atomic weight; divide each quotient by the lowest, and prefix to the result the symbol of the element.

Ex. Find the formula of a body, of which the percentage composition is Calcium 38.72; Phosphorus 20; Oxygen 41.28.

$$\text{Relative number of atoms of Ca} = \frac{38.72}{40} = .968.$$

$$\text{P} = \frac{20}{31} = .645.$$

$$\text{O} = \frac{41.28}{16} = 2.58.$$

Reducing these numbers to their simplest relation, by dividing each by the lowest, we have the number of atoms of Ca =  $\frac{.968}{.645} = 1.5$  nearly.

$$\text{P} = \frac{.645}{.645} = 1$$

$$\text{O} = \frac{2.58}{.645} = 4$$

The number for Ca not being whole, multiply each by 2, and the formula is  $3\text{Ca}_2\text{PO}_4$ , or Phosphate of Calcium.

5. Chemical combinations, and decompositions of bodies in a gaseous form, are shewn as follows :—

Ex. (1). How many cubic feet of Oxygen are requisite to consume completely 4 cubic feet of Marsh Gas, and how many cubic feet of  $\text{CO}_2$  and of  $\text{H}_2\text{O}$  will be formed?

Express the combination by an equation, thus:  $\text{CH}_4 + 2\text{O}_2 = \text{CO}_2 + 2\text{H}_2\text{O}$ .

By the law of gaseous vols., page (18), the vols. are  $2 + 4 = 2 + 4$ .

We see, therefore, that 2 cubic feet of  $\text{CH}_4$  require 4 cubic feet of O, and that 2 cubic feet of  $\text{CO}_2$  and 4 cubic feet of water-gas are formed.

$\therefore$  4 cubic feet of  $\text{CH}_4$  require 8 of O; and 4 cubic feet of  $\text{CO}_2$  and 8 of steam are formed.

#### GENERAL QUESTIONS ON THE ALKALIS AND ALKALINE EARTHS.

1. Name the Alkaline Earths, and give their principal chemical characters. (C.)

2. How is Ammonia-gas usually prepared in the laboratory? Express by symbols the reactions that occur. What weight of Sal-ammoniac must be taken to produce 100 grammes of Ammonia-gas? (C.)

3. How is Sodium Carbonate manufactured for the purposes of commerce? Explain the process, using symbols. How many tons of pure Sodium Carbonate can be, according to theory, obtained from 100 tons of common salt? (C.)

4. Compare the compounds of Ammonia, Potash, and Soda, pointing out the several characters in which they resemble one another, and those in which they differ. (C.)

5. What weight of Sodium Carbonate must be dissolved in a litre of water, that 2000 parts by weight of the solution may contain  $\text{Na}_2\text{CO}_3$ , or one molecule of the salt? (O.)

6. Who discovered Sodium, and by what means? Describe that substance. (C.)

7. Describe and explain the preparation of a solution of Caustic Soda. (C.)

8. Explain the production of the "fur" in tea-kettles in which hard water is boiled. (C.)

9. Explain the difference between Limestone, Quicklime, and Slaked Lime. (O.)

10. The analysis of a mineral gives Silica = 55, Lime = 25.6, magnesia = 18.3,—find its formula:  $\text{SiO}_2 = 60$ ;  $\text{CaO} = 56$ ;  $\text{MgO} = 40$ . (C.)

11. How is Hyposulphite of Soda prepared? For what purpose is it chiefly used? (C.)

12. What is the composition of Ammonia—(1) by weight, (2) by volume. (C.)

13. How is Potassium obtained, and what are its properties? (O.)

14. How is Lime obtained? What are its chemical properties? (O.)

15. What weights of Sulphuric, Nitric, and Hydrochloric Acids, respectively, are required to neutralize 212 parts by weight of Carbonate of Sodium? What is the weight of the solid product in each case? (O.)

16. What relative weights of Carbonate of Calcium, Nitrate of Ammonium, and Chlorate of Potassium must be taken to produce equal volumes of Carbonic Acid, Nitrous Oxide, and Oxygen? ( $\text{Ca} = 20$ ,  $\text{Cl} = 35.5$ ,  $\text{K} = 39.1$ ). (O.)

17. What is the action of heat on red-chromate of Potassium? \*

18. Give a process by which Barium Salts may be distinguished from Salts of Calcium. \*

19. An analysis of a salt gave the following numbers: what is its empirical formula?

Sulphur	22.53	*
Sodium	32.39	
Oxygen	45.08	

---

100.00

---

20. A mineral has been found to contain, in 100 parts, Silicium 20.28 Magnesium 34.16, Oxygen 45.56. What is its formula? \*

### III. METALS OF THE EARTHS.

#### ALUMINUM.

*Symbol* Al. *Combining weight* 27.5. *Sp. gr.* 2.6.

This metal may be obtained from its chloride ( $\text{Al}_2\text{Cl}_6$ ) in the same manner as magnesium. It is of a white colour, malleable, and ductile, and resembles zinc both in colour and hardness. It is very sonorous. On account of its lightness and lustre it is used in the manufacture of ornaments and trinkets. An alloy, composed of 90 parts of copper and 10 of aluminum, resembles gold in appearance, and is largely manufactured under the name of **aluminum bronze**, and Abyssinian gold. When heated in a current of oxygen, aluminum burns with a brilliant bluish-white light, producing its oxide—alumina.

**Alumina** ( $\text{Al}_2\text{O}_3$ ) is the only known oxide of this metal. It occurs abundantly in combination with silicon and oxygen in felspar, also in all varieties of clay, and in every kind of soil. It is found native, almost pure and crystalline, as the **sapphire**, **ruby**, **corundum**, and **emery**.

Alumina may be prepared by adding to a solution of alum, aqueous ammonia, when the hydrate ( $\text{Al}_2\text{O}_3 + 3\text{H}_2\text{O}$ ) is thrown down as a white, gelatinous substance. The hydrate, on being

heated, parts with its water, and forms a white soft powder of pure alumina. The hydrate is readily soluble in acids, and in the caustic alkalis; but, anhydrous, it is scarcely acted on by acids. In either state, it is insoluble in water; it has no taste, and does not change colouring matter. It acts as a weak base. Alumina and its compounds are of great use to the dyer and calico-printer. Hydrated alumina has the property of forming insoluble compounds, called lakes, with colouring matter. Fast colours are given to cloth by soaking it first with a preparation of alumina, and then dipping it in the colouring matter. The alumina adheres to the fibre of the cloth as well as to the colouring matter, so that the latter cannot be washed out.

#### COMPOUNDS OF ALUMINUM.

The CHLORIDE OF ALUMINUM ( $\text{Al}_2\text{Cl}_3$ ) is one of interest, as being one of the sources of the metal. It is obtained as a yellow, anhydrous, volatile substance, by heating a mixture of alumina and charcoal in dry chlorine gas.

SULPHATE OF ALUMINUM ( $\text{Al}_2\text{SO}_4 + 18\text{H}_2\text{O}$ ) is obtained by dissolving alumina in sulphuric acid. It is made on a large scale under the name of alum cake, by the action of sulphuric acid upon clay. It is much used by the dyer and calico printer.

ALUM, the Sulphate of Aluminum and Potassium ( $\text{KAl}_2\text{SO}_4 + 12\text{H}_2\text{O}$ ), is the most valuable salt of aluminum. It may be prepared by mixing solutions of the two sulphates, and allowing the compound salt to crystallize. Alum is manufactured on the large scale by decomposing shale or clay containing iron pyrites ( $\text{FeS}_2$ ), which is found abundantly near Whitby and Glasgow, and called alum schist, or alum ore. On roasting the shale, oxygen is absorbed from the air, and sulphuric acid is produced, which, uniting with the alumina of the clay, forms sulphate of aluminum. Sulphate of iron, also formed in the

same manner, remains in solution with the sulphate of aluminum; but it is decomposed on the addition of chloride of potassium, and alum crystallizes out.

Ordinary alum has a sour-sweetish, astringent taste; it is soluble in 18 parts of cold water, and the solution has a strongly acid reaction, and dissolves iron and zinc. Alum is employed in medicine, and for dyeing.

**SILICATES OF ALUMINUM.** These are numerous and important compounds. The various clays, formed by the crumbling away of rocks, consist chiefly of these compounds. Clay is really the result of the action of air and water upon rocks containing felspar. Clay is the chief constituent of earthenware, porcelain, or china.

**Common earthenware** is clay dried in the air, and baked in a kiln. It is firm, solid, porous, brittle, and gives out a dull sound when struck. To make earthenware fit to contain liquids, it is glazed by coating it with a fusible glass, and then heating it in a kiln till the surface is covered with a glassy varnish. For the commoner kinds of earthenware, common salt is used for a glaze.

**Porcelain or china** is made of the finest kind of clay, or china clay, and powdered glass. The latter as it melts binds the clay together, and produces a mass more or less transparent. There are various kinds of porcelain, according to the recipes of the makers.

Porcelain differs from earthenware in breaking with a glassy fracture, in possessing a ringing sound when struck, and in being semi-transparent. Like earthenware, it must be glazed to make it fit to hold liquids.

#### TESTS FOR ALUMINUM COMPOUNDS.

1. Sulphide of ammonium gives, in a solution of aluminum compounds, a white precipitate of the hydrate of alumina,

with evolution of sulphuretted hydrogen. Only one other metal—zinc, gives a white precipitate with this test.

2. Caustic potash produces a white precipitate of hydrated alumina, soluble in excess. In this solution chloride of ammonium gives a white precipitate of hydrated alumina, while in a similar solution, hydrosulphuric acid gives no precipitate. This test distinguishes aluminum from zinc.

With the blow-pipe a salt of aluminum, moistened with nitrate of cobalt, on charcoal acquires a pale blue colour.

The alkalis and alkaline earths are distinguished from the other metals by giving no precipitate with sulphide of ammonium.

Glucinum, Yttrium, Erbium, Terbium, Cerium, Lanthanum, and Didymium are rare metals, prepared in the same way as aluminum, and forming compounds like it.

---

#### IV. METALS PROPER.

Div. I. *Metals not precipitable in acid solutions by Sulphuretted Hydrogen*:—Zinc, Iron, Manganese, Cobalt, Nickel, Chromium.

##### (1) ZINC.

*Symbol Zn. Combining weight 65. Sp. gr. 6·8 to 7·1.*

Zinc is found abundantly as the carbonate, called **calamine stone**; as the sulphide, called **zinc blende**, or **black-jack**; and as the red oxide.

The metal is extracted by roasting the powdered ores. During this process, the sulphur and the carbon burn away as sulphurous acid and carbonic acid respectively, and the zinc becomes oxidized. The roasted ore is then mixed with fine coal or charcoal, heated in retorts of peculiar shape, and metallic zinc is distilled over, and received in water.



**Properties.** Zinc is a bluish-white metal of crystalline structure. It is brittle at ordinary temperatures, but when heated to about  $250^{\circ}$  F. it becomes ductile and malleable. Heated still higher it becomes brittle again. It melts at about  $773^{\circ}$  F.; at a red heat it rises in vapour, and if this be exposed to the air, it burns with a yellow-green flame, forming oxide of zinc. It tarnishes in a moist atmosphere, becoming coated with a thin film of the oxide. It dissolves in dilute acids, giving off hydrogen.

Zinc, on account of its durability, cheapness, and lightness, is used instead of lead in roofing. It is also used in the construction of galvanic batteries. Iron covered with zinc is preserved from oxidation, and is said to be **galvanized**.

**Brass** is an alloy of two parts of copper and one of zinc. **German silver** is an alloy of zinc, copper, and nickel.

**OXIDE OF ZINC ( $\text{ZnO}$ ).** This is the only known oxide. It is obtained by burning the metal in air, or by precipitating a soluble salt of zinc by means of an alkali, and heating the precipitate. It is a white, insoluble powder. When heated it becomes yellow, but on cooling it returns to its original white colour. The oxide is soluble in acids.

**SALTS OF ZINC.** The most important are :—

**Sulphide of Zinc ( $\text{ZnS}$ ), or Blende,** one of the most abundant zinc ores, generally of a black colour, from its admixture with sulphide of iron, hence called **black-jack**.

**Chloride of Zinc ( $\text{ZnCl}_2$ ),** a white, soluble, very deliquescent substance, obtained by heating zinc in chlorine gas, or by dissolving the metal in hydrochloric acid. A solution of this salt under the name of **Burnett's disinfecting fluid** is largely used as an antiseptic and disinfectant, and as a preservative of wood and vegetable fibre against decay.

**Sulphate of Zinc ( $\text{ZnSO}_4 + 7\text{H}_2\text{O}$ ),** a soluble salt, the **white**

**vitriol** of commerce, obtained by dissolving zinc in sulphuric acid. It is used in medicine, and in calico printing.

**Carbonate of Zinc** ( $\text{ZnCO}_3$ ), an insoluble compound, found native, both massive and crystallized, as calamine. It is usually of a greyish or yellowish colour.

#### TESTS FOR ZINC.

1. Sulphuretted hydrogen gives no precipitate in acid solutions.

2. Sulphide of ammonium gives a **white** precipitate of hydrated sulphide of zinc.

3. Caustic potash gives a **white** precipitate of hydrated oxide of zinc, soluble in excess. Tests 2 and 3 have the same reaction with aluminum compounds. In the potash solution, sulphuretted hydrogen gives a **white precipitate** of the hydrated sulphide of zinc.

**Blow-pipe tests.** 1. A salt of zinc, mixed with carbonate of sodium on charcoal, and exposed to the reducing flame, yields a **yellow incrustation**, which becomes white on cooling.

2. Moistened with nitrate of cobalt, and heated on charcoal in the outer flame, a salt of zinc acquires a **green** colour.

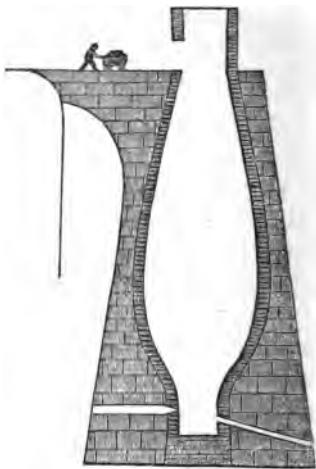
#### (2) IRON (Ferrum).

*Symbol Fe. Combining weight 56. Sp. gr. 7.8.*

Iron is more widely diffused than any other metal. It rarely occurs native. Its chief ores are **magnetic iron ore**, the **black oxide** ( $\text{Fe}_3\text{O}_4$ ); **specular iron ore**, or red hæmatite ( $\text{Fe}_2\text{O}_3$ ); **spathic iron ore**, or carbonate of iron ( $\text{FeCO}_3$ ): but the chief source of iron in this country is **clay ironstone**, an impure carbonate of iron containing about 30 to 33 per cent. of metallic iron, mingled with clay, oxide of manganese, lime, and magnesia.

Pure metallic iron is found in meteoric stones.

**Extraction of Iron.** Clay ironstone is first roasted, to drive away water, carbonic acid, and volatile substances. The roasted ore is then placed in the **blast furnace**, which is a large conical structure about 50 feet high. The ore is introduced near the top of the furnace, along with coal and limestone. At the bottom there are blast pipes, through which heated air is forced, to keep up a high temperature within the furnace. During the process of **smelting**, the limestone separates the clay (silicate of aluminum) from the oxide of iron, forming a glassy compound of silicate of calcium and aluminum, called **slag**: the oxide of iron parts with its oxygen to the coal, and falls to the bottom of the furnace, as liquid iron. The liquid slag and iron are drawn off at intervals,—the latter running into channels of sand, and is known as **pig**, or **cast iron**.



Cast iron is not pure iron. It contains about five per cent. of carbon, and small quantities of silicon, sulphur, manganese, phosphorus, &c.; hence there are several varieties of cast iron. Pig iron is highly crystalline and brittle.

**Wrought or Malleable Iron** is formed from cast iron by removing its impurities. This is done by melting it in a furnace of the shape of a baker's oven. During the process of **puddling**, as it is called, the carbon, &c., is expelled, and the melted iron thickens, and is rolled into lumps. These lumps are taken out and hammered to expel the liquid slag, and any remaining liquid cast iron, and, at the same time, coherence is given to the metal. The iron is again melted and rolled into

bars, to make the fibres parallel, and thus greater toughness is acquired.

Wrought iron is not quite pure : it contains about  $\frac{1}{300}$  of its weight of carbon, and about  $\frac{1}{300}$  of silicon. It is fibrous in texture, and very tough ; hence it is used for all purposes where a great strain is required.

Wrought iron, when exposed to continued vibration and concussion, loses its fibrous texture and becomes crystalline, and consequently brittle. This change accounts for the breaking of railway-carriage axles, &c.

Steel is made from wrought iron by adding carbon to the latter. Bars of wrought iron and charcoal powder are placed in iron boxes, and exposed to a red heat for 40 or 50 hours. At the end of that time, the iron will have taken up from 1·3 to 1·5 per cent. of carbon. It is now more fusible than before ; it has lost its fibrous texture, and has become fine grained.

The most important character of steel is its power of assuming great hardness and elasticity when heated to redness and suddenly cooled. By varying the heat and the time of cooling, different degrees of hardness and brittleness are given to steel. This process is called **tempering**.

BESSEMER'S process of manufacturing steel is much quicker and cheaper. Atmospheric air is sent through fused cast iron, which burns away nearly all the carbon and silica. To the wrought iron thus made, pure molten cast iron is added in sufficient quantity to convert the whole into steel. In this way six tons of cast iron may be converted into steel in twenty minutes.

**Properties.** Pure iron has a white colour, and a lustre almost equal to silver. It is softer than bar iron, but very tough. In bar iron the colour is bluish white or grey. It is the most tenacious of all metals. It is strongly magnetic up to a red heat. Iron does not oxidize or tarnish in dry air ; heated to redness it becomes covered with a scaly coating of black oxide, and at a white heat it burns brilliantly, producing

the same oxide. Polished iron does not tarnish in pure water, but if carbonic acid be present, and air find access, the iron rusts, forming the hydrated sesquioxide ( $2\text{Fe}_2\text{O}_3 + 3\text{H}_2\text{O}$ ). Iron at a red heat decomposes water, liberating hydrogen, and forming the black oxide. It is soluble in dilute sulphuric and hydrochloric acids, with evolution of hydrogen.

**OXIDES OF IRON.** Of these there are four, viz :—

- (1) **Protoxide, or Ferrous Oxide** ( $\text{FeO}$ ), the base of the green or ferrous salts.
- (2) **Sesquioxide, or Ferric Oxide** ( $\text{Fe}_2\text{O}_3$ ), the base of the red or ferric salts.
- (3) **Black or Magnetic Oxide** ( $\text{Fe}_3\text{O}_4$ ), which does not form any definite salts.
- (4) **Ferric Acid** ( $\text{FeO}_3 + \text{H}_2\text{O}$ ), a weak acid, forming salts with potassium.

1. **FERROUS OXIDE** has a great affinity for oxygen, passing into a higher oxide, hence the difficulty of obtaining it pure. Its hydrate is obtained by adding caustic potash to a solution of a ferrous salt, when a **white** precipitate is thrown down, which changes to a **greenish-brown** colour if oxygen be present. This oxide colours glass green; and all its salts, when pure, are of a pale green colour.

The most important salts of this oxide are—

**Ferrous Sulphide, or Sulphuret of Iron** ( $\text{FeS}$ ), prepared by bringing a bar of iron, heated to whiteness, in contact with roll sulphur. It is largely used in the preparation of sulphuretted hydrogen.

**Bisulphide of Iron** ( $\text{FeS}_2$ ) is found abundantly as **iron pyrites**; it is used in the production of sulphuric acid.

**Ferrous Chloride, or Protochloride of Iron** ( $\text{FeCl}_2$ ), a soluble salt, formed by passing dry hydrochloric acid gas over metallic iron, or by dissolving iron in hydrochloric acid.

**Ferrous Sulphate**, or **Protosulphate of Iron** ( $\text{FeSO}_4 + 7\text{H}_2\text{O}$ ), a soluble salt, commonly called **green vitriol** and **copperas**. It is prepared by dissolving pure iron, or ferrous sulphide, in oil of vitriol. On evaporation large green crystals are formed, which are used in making black dyes and writing ink. This salt has great affinity for oxygen.

**Ferrous Carbonate**, or **Protocarbonate of Iron** ( $\text{FeCO}_3$ ), an insoluble salt; is found native abundantly as clay ironstone. It is contained in all ferruginous springs. When the carbonic acid which held it in solution escapes, hydrated ferric oxide is deposited, hence those ochrey deposits around such springs.

2. **FERRIC OXIDE**, the **Sesquioxide—Peroxide**, or **Red Oxide** ( $\text{Fe}_2\text{O}_3$ )—occurs native in great abundance as one of the ores of iron. It is obtained **anhydrous** by heating green vitriol; in this state it is of a red colour, **rouge** being one of its varieties. The **hydrate**, obtained by precipitating it from a solution of one of the ferric salts by means of caustic potash, is of a **brown** colour.

The salts of this oxide are all of a **brown** colour. The chief of these are—

**Ferric Chloride**, or **Perchloride of Iron** ( $\text{Fe}_2\text{Cl}_6$ ) obtained in anhydrous reddish-brown scales by passing dry chlorine gas over red-hot iron. The hydrate is obtained in red crystals by dissolving ferric oxide in hydrochloric acid.

**Ferric Sulphate**, or **Persulphate of Iron** ( $\text{Fe}_2(\text{SO}_4)_3$ ), obtained by the action of strong sulphuric acid.

3. **BLACK OR MAGNETIC OXIDE OF IRON** ( $\text{Fe}_3\text{O}_4$ ) is found native as the mineral **loadstone**; it is black in colour, and has a metallic lustre.

This is the oxide formed by heating iron in air or oxygen.

4. **FERRIC ACID** is unimportant.

## TESTS FOR IRON.

Iron forms two classes of salts, **protosalts** and **persalts**, which are readily distinguished from each other by the colour of their solutions. The **proto-salts** in solution and crystallized have a **pale green** colour; the **per-salts** have a **yellow or reddish-brown** colour.

## TESTS FOR A PROTOSALT OF IRON.

1. Sulphuretted hydrogen in acid solutions gives **no precip.**
2. Sulphide of ammonium gives a **black** precip. of the protosulphide.
3. Caustic potash gives a **whitish-green** precip. of the hydrated protoxide, passing quickly into **brown** on exposure to the air.
4. Ferrocyanide of potassium gives a **pale blue** precip., deepening in colour on exposure.

## TESTS FOR A PERSALT.

1. Sulphuretted hydrogen gives a **white** deposit of sulphur.
2. Sulphide of ammonium gives a **black** precip.
3. Caustic potash gives a **brownish-red** precip.
4. Ferrocyanide of potassium gives a **deep blue** precip.

With the blow-pipe both classes of salts act alike. They give a **green** gloss to borax in the reducing flame, which becomes colourless or yellowish in the oxidating flame.

## (3) MANGANESE.

*Symbol* Mn. *Combining weight* 55. *Sp. gr.* 8.

The ores of this metal are abundant. The most important ore is the black oxide. The metal is not used in the arts, and its preparation is a mere matter of curiosity. It may be

obtained by heating the oxide very strongly with charcoal. It is a metal of a greyish-white colour, brittle, and hard enough to scratch glass. It oxidizes on exposure to the air, decomposes water at the ordinary temperature, and should be kept in sealed tubes, or under naphtha.

The compounds of the metal are useful in the laboratory for preparing hydrochloric acid, chlorine, and as a source of oxygen. They are also used in colouring glass, in the preparation of cast steel, and in dyeing.

**OXIDES OF MANGANESE.** The most important are the

**Protoxide, or Manganous Oxide ( $\text{MnO}$ ).**

**Sesquioxide, or Manganic Oxide ( $\text{Mn}_2\text{O}_3$ ).**

**Black, Manganic Dioxide, Bin oxide, or Peroxide ( $\text{MnO}_2$ ).**

**Manganic Acid ( $\text{MnO}_2 + \text{H}_2\text{O}$ ).**

**Permanganic Acid ( $\text{Mn}_2\text{O}_7 + \text{H}_2\text{O}$ ), or ( $\text{HMnO}_4$ ).**

The **Protoxide ( $\text{MnO}$ )**, obtained by heating the carbonate in a current of hydrogen, is an olive green powder, which absorbs oxygen from the air, and which, on ignition, changes into a higher oxide of a brown colour. Its hydrate ( $\text{MnO} + \text{H}_2\text{O}$ ) is precipitated as a whitish substance in any soluble salt by means of an alkali. The hydrate rapidly deepens in colour consequent upon absorbing oxygen. This oxide forms with acids pink-coloured salts.

The **Sesquioxide ( $\text{Mn}_2\text{O}_3$ )** occurs anhydrous in nature as braunite, and in a hydrated state in manganite. This is the oxide into which the protoxide changes on exposure to air. It is a weaker base than the corresponding oxide of iron, or alumina. The violet colour given to glass, and the colour of the amethyst are due to this oxide.

The **Black Oxide ( $\text{MnO}_2$ )** is the most abundant ore of manganese. When heated, it gives off one-third of its oxygen; if heated with concentrated sulphuric acid, it expels one-half its



oxygen. It is largely used in the preparation of chlorine and bleaching powder.

**Manganic Acid** ( $\text{MnO}_2 + \text{H}_2\text{O}$ ) has never been obtained anhydrous. It is always found combined with water, or a metallic oxide. A compound of this acid and potassium is obtained by melting equal parts of caustic potash and black oxide together, when a green mass, **manganate of potassium** ( $\text{K}_2\text{MnO}_4$ ), is formed. If this be dissolved in a little water, a green coloured solution is obtained, which changes into bright red and purple on the addition of water. On account of these changes, this substance is called **mineral chameleon**.

**Permanganic Acid** ( $\text{HMnO}_4$ ), like the manganic, is never found anhydrous. The purple liquid obtained by dissolving **mineral chameleon** in water is a solution of **permanganate of potassium**.

The manganates and permanganates part with some of their oxygen in the presence of organic matter. Under the name of **Condy's liquids**, they are used as disinfectants.

**SALTS OF MANGANESE.** The two most important are the **chloride** and the **sulphate**.

The **Chloride of Manganese** ( $\text{MnCl}_2 + 4\text{H}_2\text{O}$ ) is the residue left in the preparation of chlorine from the **black oxide** and hydrochloric acid. Iron is usually present; but the pure chloride, when crystallized by evaporation, is of a delicate pink colour.

The **Sulphate of Manganese** ( $\text{MnSO}_4 + 5\text{H}_2\text{O}$ ) is obtained by the action of sulphuric acid on the black oxide. It is a very soluble salt, of a beautiful rose colour, largely used by the dyer to give a permanent brown dye.

#### TESTS FOR MANGANESE.

1. Sulphuretted hydrogen gives no precipitate.
2. Sulphide of ammonium gives a **flesh-coloured** precipitate

of hydrated sulphide of manganese, which becomes brown on exposure to air.

3. Caustic potash gives a white precip. of hydrated protoxide, which changes to brown on exposure to air.

**Blow-pipe tests.** 1. A compound of manganese, melted with borax in a loop of platinum wire, gives in the outer flame a violet-coloured bead, which becomes colourless in the inner flame.

2. Fused on platinum foil with carbonate of soda, a bluish-green bead is produced.

#### (4) COBALT.

*Symbol* Co. *Combining weight* 59. *Sp. gr.* 8.95.

This metal is commonly found in combination with arsenic and sulphur. It is not easy to obtain it pure, as its ores are very complicated. It is of a reddish-grey colour, hard, strongly magnetic, and very tenacious. It is feebly acted upon by hydrochloric and sulphuric acids; and it slowly oxidizes on exposure to the air. Cobalt in the metallic state is not used in the arts; many of its compounds are remarkable for the beauty and brilliancy of their colour. They are used for pigments, and give to glass a beautiful blue tint.

**OXIDES OF COBALT.** There are two, the protoxide ( $\text{CoO}$ ) and the sesquioxide ( $\text{Co}_2\text{O}_3$ ). The former is a brown powder, prepared by carefully heating the hydrate, precipitated in a solution of one of the cobalt salts on the addition of an alkali. It is soluble in acids, and forms solutions of a blue colour when concentrated, but become pink on dilution. It is largely used for colouring porcelain.

The sesquioxide forms no salts; it is a black, insoluble powder.

**SALTS OF COBALT.** There are several of these, but the

most important is the chloride ( $\text{CoCl}_2$ ). This is obtained hydrous, as ruby-red crystals, by dissolving the protoxide in hydrochloric acid; when anhydrous, the crystals are blue. A weak solution of this salt is of a pink colour, and is used as a sympathetic ink. The writing, dried in the air, is invisible, but becomes blue when heated. On exposure to moisture, the characters again become invisible.

The Nitrate of Cobalt ( $\text{Co}_2\text{NO}_3 + 6\text{H}_2\text{O}$ ), prepared by dissolving the oxide in nitric acid, is useful in blow-pipe testing. In the outer blow-pipe flame this salt yields, with alumina compounds, a blue colour; with zinc compounds, green; with those of magnesium, a pale pink.

#### TESTS FOR COBALT.

1. Caustic potash gives a blue precip., which becomes rose-coloured in excess.

2. Before the blow-pipe, the borax bead acquires a deep blue colour.

#### (5) NICKEL.

*Symbol Ni. Combining weight 59. Sp. gr. 8.8.*

This metal is always found associated with cobalt. Its most abundant ores are *kupfernickel* (arsenide of nickel), and *speiss* (arseniosulphide of nickel). The process of extraction of the metal from its ores is complicated. It is prepared chiefly for the manufacture of German silver, which is an alloy of nickel, zinc, and copper.

Pure nickel is a brilliant, silver-white, hard, ductile metal strongly magnetic—a property which it loses if heated beyond  $626^\circ\text{F}$ . It becomes oxidized if exposed to air of a high temperature. Sulphuric and hydrochloric acids dissolve it, but slowly.

**OXIDES OF NICKEL.** There are two oxides, the protoxide ( $\text{NiO}$ ), and the sesquioxide ( $\text{Ni}_2\text{O}_3$ ); only the first forms salts.

The protoxide may be obtained anhydrous by heating the carbonate or nitrate of metal in a covered crucible; or as a hydrate, by adding caustic potash to a solution of a nickel salt. The oxide has a greenish colour, and the nickel salts it forms have a delicate green colour.

**SALTS OF NICKEL.** Of these the most important are the sulphate ( $\text{NiSO}_4 + 7\text{H}_2\text{O}$ ), and the chloride ( $\text{NiCl}_2$ ).

#### TESTS FOR NICKEL.

1. Sulphuretted hydrogen gives no precipitate.
2. Sulphide of ammonium gives a black precip.
3. Caustic potash gives a pale green precip. (hydrated oxide of nickel), insoluble in excess.
4. Ammonia gives a pale green precip., soluble in excess, forming a blue solution.

**Blow-pipe tests.** Nickel salts give with borax in the outer flame a reddish-yellow glass; and in the inner flame a dark grey opaque bead.

### (6) CHROMIUM.

*Symbol Cr. Combining weight 52.5. Sp. gr. 6.8.*

Chromium is found in combination with iron as chrome ironstone, chiefly in North America and Sweden; and also combined with lead, as chromate of lead—a very beautiful mineral, from which the metal was first obtained in 1797.

Chromium is not used in the arts, but its compounds are much used by the painter and dyer.

Pure chromium is the most infusible of all the metals, and is scarcely attacked by the strongest acids.

**OXIDES OF CHROMIUM.** There are four of these, closely corresponding to those of iron, viz., a protoxide ( $\text{CrO}$ ), a sesquioxide ( $\text{Cr}_2\text{O}_3$ ), a chromo-chromic oxide ( $\text{CrO}, \text{Cr}_2\text{O}_3$ ), and

**chromic anhydride**, or **chromic acid** ( $\text{CrO}_3$ ). Of these the most important are the **sesquioxide** and **chromic anhydride**.

The **Sesquioxide** or **Chromic Oxide** ( $\text{Cr}_2\text{O}_3$ ) is obtained anhydrous, of a fine green colour, by heating to redness the chromate of mercury. It is obtained as a greyish-green hydrate, by boiling alcohol with a solution of bichromate of potassium, acidulated with sulphuric acid. This oxide forms soluble salts of a green or violet colour, sweet to the taste, and poisonous.

The colouring ingredient of greenstone, the emerald, and other minerals, consists of chromic oxide.

**Chromic Anhydride—Chromic Acid**, ( $\text{CrO}_3$ ) is easily obtained by mixing equal parts of a solution of bichromate of potassium and oil of vitriol, when chromic anhydride separates in crimson-red prisms. It is very soluble in water, and the solution then has an acid reaction, forming, with bases, salts called **chromates**. Organic matter deprives it of half its oxygen, and reduces it to the sesquioxide. To this acid the ruby owes its colour.

THE SALTS OF CHROMIUM are numerous: the most important soluble salts are the **chromate** and **bichromate** of potassium.

**Chromate of Potassium** ( $\text{K CrO}_4$ ) is the source of all the preparations of chromium. It is obtained by heating chrome ironstone in a furnace along with nitrate or carbonate of potassium. The product, digested in water, gives a yellow solution, which, on evaporation, deposit yellow crystals.

**Bichromate of Potassium** ( $\text{K}_2\text{Cr}_2\text{O}_7$ ) is prepared by adding to a solution of the **chromate** a little sulphuric acid. One half of the potassium is removed by the acid, and the other half remains in combination with two atoms of chromic acid; thus the **chromate** is converted into the **bichromate**. This salt

crystallizes in beautiful red crystals. It is largely manufactured for the preparation of chrome pigments.

**Chromate of Lead** ( $\text{PbCrO}_4$ ) is an insoluble salt, prepared by precipitating a solution of acetate of lead by bichromate of potassium. It is the **chrome yellow** of the painter. From this a **bichromate of lead** of a splendid scarlet colour may be obtained by boiling the **chromate** with lime water.

**Chromate of Silver** ( $\text{Ag}_2\text{CrO}_4$ ) is an insoluble salt of a ruby-red colour when crystallized. It is prepared by mixing solutions of nitrate of silver and bichromate of potassium.

Chromium forms compounds with chlorine, of which a very interesting one is **chloro-chromic acid** ( $\text{CrO}_3 + \text{Cl}_2$ ). It is prepared by fusing 10 parts of common salt with 17 parts of chromate of potassium, and then gently heating in a retort with 30 parts of oil of vitriol. It distils over as a deep red vapour, condensing to a dense red liquid, much like bromine. The vapour is more irritating than chlorine; if the liquid be dropped into a strong solution of ammonia or alcohol, it takes fire.

#### TESTS FOR CHROMIUM COMPOUNDS.

1. Caustic potash gives a green precip., which dissolves with a green colour in excess in salts of the sesquioxide.

2. The **chromates** in solution give, with a salt of lead a **yellow** precip.; with nitrate of silver a **red** precip.; with nitrate of mercury an **orange** precip.

**Blow-pipe test.** Compounds of chromium fused with borax give an **emerald green** bead; fused on the platinum foil with nitrate of potash, they form yellow chromate of potash.

**DIV. II.** *Metals precipitated by Sulphuretted Hydrogen in neutral, acid or alkaline solutions, viz. :—Copper, Lead, Silver, Mercury, Bismuth, and Cadmium.*

**(1) COPPER (Cuprum).**

*Symbol Cu. Combining weight 63.5. Sp. gr. 8.93.*

This useful metal is found native in North America, and in the state of ore in Cornwall, Chili, South Australia, and Spain. The most common ore is copper pyrites, a compound of sulphur, copper, and iron ( $\text{Cu}_2\text{S}, \text{Fe}_2\text{S}_3$ ). Next in abundance to this is the sulphide of copper ( $\text{Cu}_2\text{S}$ ), and the less abundant are the carbonates of copper or malachite, and the red and black oxides.

Copper is extracted on the large scale from the carbonate or oxide, by reducing the ores with carbon and silica in a furnace. The process of extraction from copper pyrites is very complicated.

The chief seat of copper smelting is Swansea.

**Properties.** Copper is the only red metal. It is rather hard, very tenacious, ductile, and malleable, and is one of the best conductors of heat and electricity. It does not oxidize in air at ordinary temperatures, but if heated to redness a layer of oxide is formed upon the surface. Pure water does not affect it; but sea water and solutions of chlorides corrode it. Nitric acid dissolves it readily. Oil of vitriol in the cold does not affect it, but if heated along with it, it is decomposed. Hydrochloric acid dissolves it when finely divided.

It forms very useful alloys. With zinc it forms brass; with tin it forms bronze, bell-metal, gun-metal, and speculum metal.

**OXIDES OF COPPER.** There are two well defined oxides, which are found native, viz., the red or cupreous oxide ( $\text{Cu}_2\text{O}$ ) and the black or cupric oxide ( $\text{CuO}$ ).

**Cupreous Oxide, the Red or Suboxide of Copper ( $\text{Cu}_2\text{O}$ )**

occurs native in ruby-red crystals. It may be prepared by heating together equal parts of copper filings and the black oxide of copper in a covered crucible, or by boiling a solution of sulphate of copper with sugar and caustic potash. This oxide forms colourless salts with acids, which absorb oxygen from the air, and pass into compounds of the cupric oxide. It is chiefly used to stain glass a beautiful ruby colour.

**Cupric Oxide**, or the **Black Oxide of Copper** ( $\text{CuO}$ ), is obtained by heating copper to redness in air, or by heating to redness nitrate of copper in an earthen crucible; or by precipitating the hydrate from one of its salts, by means of caustic potash. This oxide imparts to glass a beautiful green colour. It is soluble in oils and fats, therefore it is necessary to keep copper cooking vessels perfectly clean and bright. It is soluble in acids, forming salts of a green or blue colour.

**SALTS OF COPPER.** Of the numerous salts of copper the chief are—

(1). **Sulphate of Copper, Cupric Sulphate, or Blue Vitriol** ( $\text{CuSO}_4 + 5\text{H}_2\text{O}$ ), which may be obtained by boiling copper with sulphuric acid, diluted with half its quantity of water. It forms large blue crystals, soluble in water. The water of crystallization is expelled by heating the salt to  $392^\circ \text{F}$ ., when it takes the form of a white powder. Heated to bright redness, the salt is decomposed. This sulphate is largely used in dyeing, and in the preparation of pigments.

(2). **Nitrate of Copper, Cupric Nitrate** ( $\text{Cu}_2\text{NO}_3 + 6\text{H}_2\text{O}$ ), is a corrosive, soluble salt, obtained in the form of deep blue crystals by dissolving copper in nitric acid, and evaporating the solution. Heat decomposes it, expelling the nitric acid, and leaving the black oxide.

(3). **Carbonate of Copper** has never been obtained pure; it is always associated with the hydrated oxide of copper. In this form it is found as the mineral **chessylite** ( $2\text{CuCO}_3 +$



$\text{CuO}, \text{H}_2\text{O}$ ), and as the beautiful mineral **malachite** ( $\text{CuCO}_3 + \text{CuO}, \text{H}_2\text{O}$ ), which is much used for ornamental purposes.

(4). **Chloride of Copper, Cupric Chloride** ( $\text{CuCl}_2$ ), is obtained by the action of chlorine gas upon copper; or by dissolving the black oxide in hydrochloric acid, and evaporating, when hydrated crystals of a green colour are formed ( $\text{CuCl}_2 + 2\text{H}_2\text{O}$ ). It is very soluble in water and alcohol; a solution of the latter burns with a splendid green flame. All the salts of copper are very poisonous. Raw white of egg is a good antidote.

#### TESTS FOR COPPER.

As the salts of the suboxide are rarely met with, the following tests apply only to those of the black oxide:—

1. **Sulphuretted Hydrogen** gives a black precip. of sulphide of copper ( $\text{CuS}$ ).

2. **Caustic Potash** gives a pale blue precip., which becomes black on boiling.

3. **Ferrocyanide of Potassium** gives a red-brown precip., soluble in ammonia.

4. Polished iron wire, introduced into a solution of a copper salt, becomes covered with a deposit of metallic copper.

**Blow-pipe tests.** A salt of copper, heated with carbonate of soda on charcoal in the inner flame, yields metallic copper, which is easily known from its colour.

#### (2) LEAD (Plumbum).

*Symbol Pb. Combining weight 207. Sp. gr. 11.36.*

The chief ore of lead is **galena**—a sulphide of the metal and silver—which is principally found in England and Spain.

In extracting the lead from galena, a little lime is added to the ore, in order to remove any siliceous matter it may contain. The process of extraction is as follows:—The crushed ore is first roasted in a reverberatory furnace, through which a

strong current of air is allowed to pass. During this roasting, a portion of the sulphide is oxidized, and becomes a sulphate, whilst the sulphur of another portion is burnt away as sulphurous anhydride, leaving oxide of lead, and another portion remains undecomposed. When the roasting is considered complete, the furnace is closed, and the heat suddenly raised. The oxide and sulphate of lead decompose the remaining sulphide, and the whole mass is converted into metallic lead and sulphurous anhydride.

**Properties.** Lead is a bluish-white metal, very soft, highly malleable, but not very ductile. It melts at  $620^{\circ}$  F., and at a high temperature it absorbs oxygen from the air, and gives off white fumes of oxide. It slowly tarnishes on exposure to the air. **Pure water** has no action upon the metal, but water charged with air corrodes it, and the oxide of lead thus formed dissolves in the water. If there be any chlorides or nitrates present in the water, the corrosion of the metal is greatly increased. Water containing sulphates or carbonates, as spring water, form insoluble salts of lead, and, therefore, does not become contaminated; but if it contain carbonic acid in excess, a carbonate of lead becomes soluble. On account of the use of lead pipes and lead cisterns, and the poisonous nature of lead compounds, the action of water upon the metal deserves special attention.

Lead forms many useful alloys. **Small shot** is an alloy of lead and arsenic; **type metal** is an alloy of lead and antimony; **pot-metal**, an alloy of lead and copper.

**OXIDES OF LEAD.** There are four of these, viz., a **suboxide** ( $\text{Pb}_2\text{O}$ ), which is of no importance; a **protoxide** ( $\text{PbO}$ ); a **oxide** or **peroxide** ( $\text{PbO}_2$ ); and a **red oxide** ( $\text{Pb}_3\text{O}_4$ ).

The **Protoxide of Lead—Plumbic Oxide** ( $\text{PbO}$ )—is obtained by heating lead in a current of air. If heated below the point of fusion, a yellow powder is formed called **massicot**; but if the metal be fused, it forms a scaly mass of a brick-red colour,

called **litharge**. Boiled with oil, it forms sticking-plaster. This oxide is slightly soluble in water, but soluble in solutions of the caustic alkalis. It combines with bases, and forms an important class of salts. A solution of this oxide in lime-water is used as a hair-dye. It is also used in the manufacture of flint glass, and in the making of glaze for common earthenware.

**Peroxide of Lead—Plumbic Dioxide, ( $\text{PbO}_2$ )**—is usually obtained, as a brown powder, by digesting powdered red lead in boiling nitric acid, diluted with four or five times its bulk of water. It is insoluble in water and acids. When heated, it parts with half its oxygen, and becomes a protoxide. Hot hydrochloric acid acts upon it, liberating chlorine, and forming chloride of lead.

**Red Oxide—Red Lead, or Minium**—is a compound of the two preceding oxides, and may be obtained by heating **massicot** to a low redness, whilst a current of air is allowed to pass over it. Its composition is represented as  $2\text{PbO}$ ,  $\text{PbO}_2$ . It is much used in glass making, paper staining, and colouring the inferior kinds of red sealing wax. Nitric acid decomposes it, forming nitrate of lead.

#### COMPOUNDS OF LEAD.

**NITRATE OF LEAD. Plumbic Nitrate ( $\text{Pb}_2\text{NO}_3$ )** is easily obtained by dissolving lead, or the protoxide, in diluted nitric acid. Milk-white crystals are formed, which are soluble in about eight parts of cold water. Heated to redness, it decomposes, evolving oxygen and red fumes of  $\text{NO}_2$ .

**ACETATE OF LEAD, or Sugar of Lead ( $\text{Pb}_2\text{C}_2\text{H}_3\text{O}_2$ )**, much resembles loaf sugar in taste and appearance, and is prepared by dissolving **litharge** in vinegar.

**CARBONATE OF LEAD—White Lead, ( $\text{PbCO}_3$ )**—occurs native along with galena. It may be prepared by adding carbonate

of soda to a solution of the acetate or nitrate of lead. It is largely used as a white paint. Heat decomposes it. Water, unless charged with carbonic acid gas, does not dissolve it.

**SULPHIDE OF LEAD**—Galena, ( $\text{PbS}$ ) is the chief ore of the metal; it may be formed artificially by fusing lead with sulphur, or by precipitating any lead salt by sulphuretted hydrogen.

**CHLORIDE OF LEAD** ( $\text{PbCl}_2$ ) is best prepared by adding hydrochloric acid to a solution of nitrate of lead: a white precipitate, soluble in about 33 parts of boiling water, is obtained.

**IODIDE OF LEAD** ( $\text{PbI}_2$ ) is obtained as beautiful yellow spangles, by adding iodide of potassium to a solution of nitrate of lead, dissolving the precipitate in hot water, and allowing it to cool.

#### TESTS FOR LEAD.

1. Sulphuretted hydrogen, and sulphide of ammonium, give a **black precip.** of the sulphide.

2. Caustic potash gives a **white precip.** of the hydrated oxide, soluble in excess.

3. Iodide of potassium gives a **yellow precip.** of iodide of lead.

4. Bichromate of potassium gives a **yellow precip.** of chromate of lead.

**Blow-pipe test.** Salts of lead on charcoal give a soft white bead of lead, surrounded with a yellow incrustation of the oxide.

#### (3) SILVER (Argentum).

*Symbol Ag. Combining weight 108. Sp. gr. 10.53.*

Silver is found native, and in combination with sulphur, antimony, copper, iron, and gold. Galena, the sulphide of

lead, usually contains small quantities of it. The mines of Peru, Mexico, Kongsberg in Norway, Schneeberg and Freyberg in Saxony, are the chief sources of silver.

The process of extracting the metal from its ores is complicated, and differs according to the kind of ore, and the custom of the country where it is found. Silver ores are not **smelted** like those of iron, but another method, called **amalgamation**, is employed to extract the metal. In Germany, the usual method is to roast, at low red heat, the powdered ore along with salt. By this means the sulphide of silver is converted into the chloride. The mass is then put into casks made to revolve, and scraps of iron and water are added. During the agitation the iron reduces the chloride to the state of metallic silver. Mercury is then added, and the mixture again agitated, and a liquid amalgam of mercury and silver, (and gold, if any be present,) is formed. This amalgam is strained, and then heated, when the mercury distills off, and impure silver remains.

**Pure** silver may be obtained by dissolving a silver coin in nitric acid, when nitrate of silver is formed. The solution will be greenish-blue from the presence of copper. A solution of common salt is added to the nitrate, and chloride of silver is precipitated. The chloride is washed and melted in a crucible along with carbonate of sodium,—the result is pure silver and common salt.

**Properties.** Pure silver is white, and of a brilliant lustre. It is highly malleable and ductile, and is the best conductor of heat and electricity known. In hardness, it lies between copper and gold. It fuses at 1873° F. It does not oxidize in air, but if melted, it absorbs 22 times its bulk of oxygen, which, however, it expels on cooling. Cold nitric acid, and boiling oil of vitriol dissolve it readily. Silver blackens slowly in the air, on account of the presence of sulphuretted hydrogen, which readily combines with the metal, and forms the sulphide. The brilliancy of silver makes it useful for the covering of reflectors in lighthouses. Silver stains glass an orange colour.

**ALLOYS OF SILVER.** The most common of these is an alloy of silver and copper, for coinage and jewellers' work. Copper is added to give silver sufficient hardness for extensive use. In English standard silver, the amount of copper is 7·5 per cent.

**OXIDES OF SILVER.** There are three: a **suboxide** ( $\text{Ag}_2\text{O}$ ), which is a black powder, easily decomposed, and soluble in ammonia; a **protoxide** ( $\text{Ag}_2\text{O}$ ), which is the base of the silver salts; and a **peroxide**, the composition of which is not yet determined.

The **Protoxide of Silver**, or **Argentio Oxide** ( $\text{Ag}_2\text{O}$ ), is obtained as a brown powder by adding caustic potash to a solution of any silver salt. Heat and light decompose it.

#### COMPOUNDS OF SILVER.

The two most important are the **nitrate** and **chloride** of silver.

**Nitrate of Silver**, or **Argentio Nitrate** ( $\text{AgNO}_3$ ), may be obtained as crystals, by dissolving silver in nitric acid, and evaporating the solution. These are soluble in an equal weight of cold water. When fused, the nitrate is cast into moulds, and then forms the sticks of **lunar caustic**. This salt, in contact with organic matter, is blackened by sunlight. It is used in the preparation of marking ink for linen.

**Chloride of Silver** ( $\text{AgCl}$ ) is found native as **horn silver**. It is obtained as a dense curdy white mass by adding a solution of any chloride to a soluble salt of silver. When heated, it melts, and on cooling crystallizes into what is called **horn silver**. Sunlight changes it to a violet colour, but if organic matter be present, the chloride is instantly blackened. This property makes it useful in photography. This salt is insoluble in water. Hyposulphite of sodium dissolves it, and on account of this, it is used for "fixing" photographic pictures—that is, giving them a permanent character by removing the unaltered salt of silver.

The other compounds, as the Sulphide ( $\text{Ag}_2\text{S}$ ), Bromide ( $\text{AgBr}$ ), Iodide ( $\text{AgI}$ ), Fluoride ( $\text{AgF}$ ), and Sulphate ( $\text{Ag}_2\text{SO}_4$ ), do not require any special mention.

The soluble salts of silver are colourless, and act as irritant poisons.

#### TESTS FOR SILVER.

1. Sulphuretted hydrogen and sulphide of ammonium give a **black** precip. of the sulphide.

2. Caustic potash gives a **brown** hydrated oxide, insoluble in excess.

3. Hydrochloric acid gives a **white curdy** precip., insoluble in nitric acid, but **soluble** in ammonia.

**Blow-pipe test.** Heated with carbonate of soda on charcoal, a white bead of silver is obtained.

#### (4) MERCURY (Hydrargyrum).

*Symbol* Hg. *Combining weight* 200. *Sp. gr. as a liquid* 13.596.

Mercury or quicksilver is found native in small quantities, but it usually occurs as **cinnabar**, (the sulphide of mercury), an ore which is found in Spain, California, Transylvania, China, and Japan.

The metal may be extracted by simply roasting the ore, by which the sulphide is burnt off, and the mercury distilled; or by heating the ore along with slaked lime.

**Properties.** Mercury is the only metal liquid at the ordinary temperature. It has the lustre of silver. It freezes at  $37.9^\circ \text{F}$ . below zero, and is then malleable; it boils at  $662^\circ \text{F}$ ., and forms a vapour, with a sp. gr. of 6.976. It does not tarnish in air, if pure; but heated to  $700^\circ \text{F}$ . it combines with oxygen, and forms the red oxide. At ordinary temperatures it combines with chlorine, bromine, iodine, and sulphur. Nitric acid dissolves it; hydrochloric acid does not affect it; sulphuric acid, if heated with it, is decomposed.

Mercury is used in the extraction of gold and silver from their ores; in the preparation of vermilion; in medicine; and, with its amalgams, in silvering and gilding.

**OXIDES OF MERCURY.** There are two, a black suboxide ( $\text{Hg}_2\text{O}$ ), and a red oxide ( $\text{HgO}$ ), both of which form salts.

The **Suboxide of Mercury, or Mercurous Oxide** ( $\text{Hg}_2\text{O}$ ), a dark powder, is obtained from calomel (the chloride of mercury) by digesting it with caustic potash. Heat and the sun's rays convert it into the red oxide and the metal.

The **Red Oxide of Mercury, or Mercuric Oxide** ( $\text{HgO}$ ), may be obtained as a red crystalline body by gently heating the nitrate, or by heating the metal to  $700^\circ \text{F.}$ , or by adding caustic potash to a solution of **corrosive sublimate**, (bichloride of mercury,  $\text{HgCl}_2$ ). In the last method the oxide has a yellowish colour. It is slightly soluble in water; heat decomposes it into the metal and oxygen.

#### COMPOUNDS OF MERCURY.

There are two **sulphides** corresponding to the two oxides: a **subsulphide, or mercurous sulphide** ( $\text{Hg}_2\text{S}$ ), which is of little importance, and a **sulphide, or mercuric sulphide** ( $\text{HgS}$ ). This latter is **cinnibar**, the most abundant ore of mercury; it is also known as the pigment vermilion, which is prepared by heating sulphur and mercury. Mercuric sulphide is precipitated as a black body by sulphuretted hydrogen, but on drying it assumes a red colour.

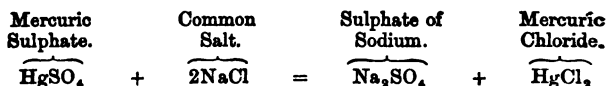
There are two **chlorides**, which are very important compounds; they are commonly known as **calomel** and **corrosive sublimate**.

The **Chloride of Mercury, Mercurous Chloride, or Calomel** ( $\text{HgCl}$ ), is obtained by precipitating a solution of nitrate of mercury by the aid of common salt, or by heating a mixture of three parts of finely divided mercury with four parts of



corrosive sublimate. Calomel is a heavy, white, tasteless powder, insoluble in water. The alkalis and lime water decompose it. Sulphuric acid has no action upon it; boiling nitric acid, and hydrochloric acid, after long boiling, decompose it. It is largely used for medicinal purposes.

The Bichloride of Mercury, Mercuric Chloride, or Corrosive Sublimate ( $\text{HgCl}_2$ ), is prepared on the large scale by heating a mixture of one part of common salt, and two and a half parts of the sulphate of the red oxide. It is also obtained by heating metallic mercury in chlorine gas. The changes in the former process are thus shown:—



Corrosive sublimate is highly poisonous; fuses at  $509^\circ \text{F.}$ ; soluble in water, alcohol, and ether. In cases of poisoning, white of egg is the usual antidote, as corrosive sublimate forms with this an insoluble compound.

Mercury forms with iodine **three iodides**, and by heating the metal along with oil of vitriol, sulphate of mercury, or mercuric sulphate ( $\text{HgSO}_4$ ) is obtained.

**Nitrates of mercury**, corresponding to the two oxides, are prepared by the action of dilute nitric acid upon the metal, and by dissolving mercuric oxide in nitric acid.

#### TESTS FOR MERCURY.

*In a solution of a salt of the suboxide:—*

1. Sulphuretted hydrogen gives a **black** precip. of subsulphide of mercury.
2. Hydrochloric acid gives a **white** precip. of calomel, soluble in hot nitric acid.
3. Caustic potash gives a **black** precip. of the hydrated suboxide.

4. Iodide of potassium gives a **green** precip. of the iodide of the suboxide.

*In a solution of a salt of the red oxide :—*

1. Sulphuretted hydrogen gives a **dirty white** precip. of the sulphide, changing through red to black.

2. Hydrochloric acid gives no precip.

3. Caustic potash gives a **bright yellow** precip. of hydrated red oxide.

4. Iodide of potassium gives a **brilliant scarlet** precip. of the iodide.

Salts of mercury in solution give a white metallic coating to strips of copper, which disappear on heating. Heated with carbonate of sodium in a tube, they deposit metallic globules.

### (5) BISMUTH.

*Symbol Bi. Combining weight 210. Sp. gr. 9.79.*

Bismuth is not an abundant metal ; it is found native, and also combined with sulphur. It is easily reduced by heating in iron tubes.

It is a hard, brittle metal, of a pinkish-white colour, and fuses at 507° F. It does not oxidize in dry air at the ordinary temperature, but if heated strongly it burns with a blue form, forming an oxide. When thrown in a powdered state into chlorine gas, it takes fire. Nitric acid dissolves it rapidly. It is chiefly used in making the alloy known as **fusible metal**. Its compounds are used for pigments and in medicine.

**OXIDES OF BISMUTH** are two in number,—a **sesquioxide**, or bismuth oxide ( $\text{Bi}_2\text{O}_3$ ), which is a yellow, insoluble powder, obtained by heating the nitrate ; and a **peroxide**, or **bismuthic oxide** ( $\text{Bi}_2\text{O}_5$ ), a brown powder, obtained by dissolving the hydrated sesquioxide in potash, and transmitting chlorine gas.

The most important soluble salt of bismuth is the **nitrate**, or **bismuth nitrate** ( $\text{Bi}_2\text{NO}_3 + 5\text{H}_2\text{O}$ ), which is prepared by dissolving the metal in nitric acid. When this salt is thrown into water, a **sub-nitrate** falls as a white precipitate, whilst an acid salt remains in the liquid.

#### TESTS FOR BISMUTH.

1. Sulphuretted hydrogen gives a **black precip.** of the sulphide ( $\text{Bi}_2\text{S}_3$ ).

2. Caustic potash gives a **white precip.** insoluble in excess, which becomes **yellow** on boiling.

3. Water gives a **white precip.**

With the **blow-pipe**. Bismuth compounds, heated on charcoal with carbonate of sodium, give a metallic bead with a yellowish incrustation of the oxide.

#### (6) CADMIUM.

*Symbol* Cd. *Combining weight* 112. *Sp. gr.* 8.6.

Cadmium is usually found in the zinc ores as sulphide of cadmium. Being more volatile than zinc, cadmium distils over first in the preparation of the former metal. It resembles tin in appearance; it is malleable and ductile, and fuses at  $442^\circ \text{F}$ . It does not oxidize in air, but when heated red-hot it takes fire, depositing a brown oxide ( $\text{CdO}$ ). Cold nitric acid, and hot sulphuric or hydrochloric acid, dissolve it readily.

The cadmium salts are colourless, and resemble those of zinc.

#### TESTS FOR CADMIUM.

Sulphuretted hydrogen gives a **yellow precip.** of the sulphide, insoluble in ammonia.

With the **blow-pipe**. Cadmium compounds, heated on charcoal with carbonate of soda, give a brown incrustation of oxide of cadmium.

Div. III. *Metals which give no precipitate with Sulphuretted Hydrogen in alkaline solutions:—Gold, Platinum, Tin, Antimony, Arsenic, &c.*

(1) GOLD (Aurum).

*Symbol Au. Combining weight 196.6. Sp. gr. 19.34.*

Gold is a metal widely diffused, and always found in the metallic state in the primitive rocks, and in the beds of certain rivers. The most celebrated mines are those of California, Australia, Mexico, Chili, Brazil, and Peru.

Gold which occurs in sand is extracted by the simple process of washing in a "cradle," and forms the "gold-dust" of commerce. Rocks containing gold are crushed to powder, and the metal extracted by the process of amalgamation with mercury.

**Properties.** Gold has a beautiful yellow colour, and, when pure, is as soft as lead. It is very malleable and ductile, and fuses at 2016° F. It is a good conductor of heat and electricity. It does not tarnish in air, and it resists the action of every acid, excepting **selenic**. A mixture of one part of nitric acid, and four parts of hydrochloric acid, called **aqua regia**, dissolves it; so also does any mixture which liberates chlorine.

Pure gold may be obtained by dissolving the ordinary metal in **aqua regia**, and adding sulphate of the protoxide of iron, when gold is precipitated as a brown powder, which becomes bright by friction.

Gold is extensively used for ornaments, &c., but on account of its softness, it requires to be mixed with other metals. Jewellers' gold is usually mixed with copper and silver. British **standard gold** is an alloy of eleven parts of gold to one of copper.

**OXIDES OF GOLD.** Of these there are two,—a **suboxide** ( $\text{Au}_2\text{O}$ ), and a **sesquioxide** ( $\text{Au}_2\text{O}_3$ ). Neither of these forms

salts with acids; but the latter, when hydrated, has acid properties, and is sometimes called **auric acid**. It forms with bases, salts called **aurates**.

### COMPOUNDS OF GOLD.

The most important of these are the **two chlorides**, the **protochloride** ( $\text{AuCl}$ ), and the **trichloride** ( $\text{AuCl}_3$ ).

The **Protochloride**, or **Aurous Chloride** ( $\text{AuCl}$ ) is obtained as a pale-yellow insoluble powder, by heating the trichloride to  $347^\circ \text{F.}$ , and thus expelling two-thirds of its chlorine.

**Trichloride of Gold**, or **Auric Chloride** ( $\text{AuCl}_3$ ) is the most important gold compound, and is obtained by dissolving the metal in aqua-regia, and evaporating to dryness. The trichloride is a red crystalline body, very deliquescent, and soluble in water, alcohol, and ether. It unites with metallic chlorides, and forms double salts, which are mostly of a yellow colour when in crystals, and red when deprived of water.

The remaining gold compounds are of no importance.

### TESTS FOR GOLD.

1. A solution of **protosulphate of iron** gives a **brown precip.** of metallic gold.

2. **Chloride of tin** gives a **purple precip.** called **purple of cassius**, which is the hydrated double stannate of gold and tin. ( $\text{Au}_2\text{SnO}_6 + \text{SnSnO}_3 + 4\text{H}_2\text{O}$ ).

### (2) PLATINUM.

*Symbol Pt. Combining weight 197.1. Sp. gr. 21.5.*

Platinum is a rare metal, always found native, and mixed with the metals palladium, rhodium, osmium, ruthenium, and iridium. Sometimes it is found mixed with gold, silver, copper, iron, and lead. Our chief supply comes from Mexico, Brazil, and the Ural Mountains.

The process of obtaining platinum is a complicated one. The crude metal is dissolved in aqua-regia, and precipitated in combination with several of the other metals by means of sal-ammoniac. This precipitate of **ammonic-platinic chloride** is heated, and produces a spongy mass called **spongy platinum**, which, when heated and pressed, forms a metallic mass.

A new method of extracting platinum has been lately introduced. The ore is melted in a furnace heated with the oxyhydrogen blow-pipe; most of the impurities are expelled, and an alloy of platinum, iridium, and rhodium is obtained, which is more useful than the pure metal.

**Properties.** Platinum has a white colour, and is capable of a high lustre; hence it was called by the Spaniards of South America **platina** (little silver). It does not tarnish in air, and can only be melted by the oxyhydrogen blow-pipe. It is malleable, and ductile. It resists the action of all acids; but aqua-regia dissolves it. If heated to redness with the caustic alkalis, it is oxidized, and combines with the alkaline base.

On account of its infusibility, and its power of resisting chemical agents, platinum is largely used for the construction of chemical apparatus.

**THE OXIDES OF PLATINUM** are two, viz., the **protoxide** ( $\text{PtO}$ ), a black powder, soluble in acids forming unstable salts, obtained by digesting the protochloride in solution of potash; and the **dioxide** ( $\text{PtO}_2$ ), obtained as a brown hydrate, by adding to a solution of nitrate of platinum half its quantity of caustic potash.

#### COMPOUNDS OF PLATINUM.

The most important are the **chlorides**, viz., the **protochloride**, or **platinous chloride** ( $\text{PtCl}_2$ ); and the **bichloride**, or **platinic chloride** ( $\text{PtCl}_4$ ).

The former is obtained by evaporating a solution of platinum in aqua-regia, and exposing the residue to a heat of  $455^\circ \text{F}$ . It

is a dark green, insoluble powder, and forms double salts with the chlorides of the alkalis.

The bichloride is the most important salt, and is obtained by evaporating a solution of the metal in aqua-regia. It forms an orange-coloured solution in water, and double salts with the chlorides of the alkaline metals.

#### TESTS FOR PLATINUM.

1. Sulphuretted hydrogen and sulphide of ammonium give respectively a **black** precip. of the sulphide, soluble in excess of sulphide of ammonia.

2. Caustic potash gives a **yellow** precip., soluble in excess.

#### (3) TIN (Stannum).

*Symbol* Sn. *Combining weight* 118. *Sp. gr.* 7.3.

Tin is always found in the state of ore, running in veins in the primary rocks, and is usually mixed with the sulphides of copper and iron. Its chief ore is **tinestone**, which is the dioxide of tin, and is found in Cornwall. Tin mines also occur in Malacca and Mexico.

The metal is extracted by first crushing and washing the ore, —then by roasting to expel sulphur, &c.,—and then by reduction and refining. English tin usually contains small quantities of arsenic, copper, iron, and lead.

**Properties.** Tin is, next to silver, the whitest of metals. It is soft, ductile, and malleable, but of little tenacity. It is a good conductor of heat and electricity; it fuses at 442° F. It tarnishes very slowly when exposed to air at the ordinary temperature; if exposed to great heat, it takes fire and burns with a brilliant white light, forming the dioxide. Strong hydrochloric acid, diluted nitric acid, and boiling concentrated sulphuric acid, act upon it readily.

Tin is largely used to protect the more oxidizable metals, and also for cooking utensils.

**ALLOYS OF TIN** are numerous. **Tin-plate** is an alloy of tin and iron : **Britannia metal** consists of equal parts of tin, brass, antimony, and bismuth : **Pewter** consists of four parts of tin and one of lead : **Queen's metal** is an alloy similar to Britannia metal. With copper, tin forms several alloys, as **speculum metal**, **bell metal**, **gun metal**, and **bronze**. With mercury, tin forms an amalgam for the silvering of looking-glasses.

**OXIDES OF TIN.** There are two oxides,—a protoxide, or stannous oxide ( $\text{SnO}$ ), and a dioxide, or stannic oxide ( $\text{SnO}_2$ ).

**Stannous Oxide** ( $\text{SnO}$ ) is an olive-brown powder, obtained by heating the hydrated protoxide in carbonic acid. The hydrated protoxide is obtained by adding a solution of a **stannous salt** to a strong solution of an alkaline carbonate.

The **Dioxide**, or **Stannic Oxide** ( $\text{SnO}_2$ ) occurs native as **tin-stone**, but it may be prepared artificially in two forms, having very different properties. When tin is acted upon by nitric acid, a white powder is produced, insoluble in acids, and is called **meta-stannic acid**. In the anhydrous state, this is of a buff colour, and is used for polishing plate, under the name of **putty powder**, and for the enamel of watch dials. If an alkaline solution be added to a solution of bichloride of tin, a white precipitate of stannic oxide is formed, soluble in acids ; this is distinguished from the former variety by the name of **stannic acid**.

#### COMPOUNDS OF TIN.

The most important are the chlorides, of which there are two, the chloride of tin, or **stannous chloride** ( $\text{SnCl}_2$ ), and the bichloride of tin, or **stannic chloride** ( $\text{SnCl}_4$ ). The former, in a hydrated state, is obtained by dissolving tin in hydrochloric acid ; it is used by the dyer and calico printer.



**Stannic chloride** is prepared by passing chlorine gas over melted tin; it is a thin, colourless liquid, emitting dense white fumes when exposed to the air. When mixed with water, it crystallizes. It is largely used in dyeing for fixing colours.

#### TESTS FOR TIN.

##### (I). *In a solution of a protosalt of tin:—*

1. Sulphuretted hydrogen and sulphide of ammonium give respectively a **dark brown** precip. of the hydrated sulphide of tin.
2. Caustic potash gives a **white** precip. of the hydrated stannous oxide, soluble in excess.
3. Chloride of gold gives a **purple** precip. (the purple of cassius).

##### (II). *In a solution of a persalt:—*

1. Sulphuretted hydrogen and sulphide of ammonium give respectively a **dirty yellow** precip. of the hydrated sulphide, soluble in excess.
  2. Caustic potash gives a **white** precip., soluble in excess.
- Before the blow-pipe, salts of tin form with carbonate of soda, in the inner flame, white beads of metallic tin.

#### (4) ANTIMONY (Stibium).

*Symbol* Sb. *Combining weight* 122. *Sp. gr.* 6.71.

Antimony is an abundant mineral, chiefly found in combination with sulphur, as a sesquisulphide ( $\text{Sb}_2\text{S}_3$ ). The metal is easily extracted from this ore by heating it with iron, when the sulphur and iron combine; or by mixing the ore with coal, and heating in a reverberating furnace.

**Properties.** Antimony is a bluish-white metal, very brittle, and highly crystalline. It fuses at  $1150^\circ\text{F}$ . At ordinary temperatures it does not oxidize in air; but if heated, it burns readily with a white flame, giving off white fumes of antimonious oxide. Powdered antimony takes fire when thrown into

chlorine gas. Aqua-regia dissolves it, so also does heated strong hydrochloric acid. Nitric acid and boiling sulphuric acid attack the metal.

Antimony is much used in alloys, as mentioned under the heads of Lead and Tin.

**OXIDES OF ANTIMONY.** There are three known oxides, of which only the first two are of any importance, viz,—a **sesquioxide** or **antimonious oxide** ( $\text{Sb}_2\text{O}_3$ ); **antimonic oxide** ( $\text{Sb}_2\text{O}_5$ ); and **antimonious acid** ( $\text{Sb}_2\text{O}_3$ ,  $\text{Sb}_2\text{O}_5$ ).

The **Antimonious Oxide**, or **Sesquioxide of Antimony**, called also the **Teroxide** ( $\text{Sb}_2\text{O}_3$ ), may be prepared by adding a solution of carbonate of soda to a solution of terchloride of antimony, when the oxide is precipitated as a yellow powder. When boiled with bitartrate of potassium (cream of tartar), it dissolves, and yields, on evaporation, crystals of tartar emetic. This oxide is the basis of the salts of antimony used in medicine. It dissolves in hydrochloric and tartaric acids. Nitric acid converts it into a higher oxide, and sulphuric acid forms with it an insoluble sulphate.

The **Antimonic Oxide**, also called **Antimonic Acid** ( $\text{Sb}_2\text{O}_5$ ), is obtained by the action of nitric acid upon antimony. It is a pale-yellow powder, insoluble in water, and loses part of its oxygen on the application of heat.

### COMPOUNDS OF ANTIMONY.

Antimony has a strong affinity for chlorine, and forms with it two chlorides, viz., **antimonious chlorine**, or **trichlorine of antimony** ( $\text{SbCl}_3$ ), and **antimonic chlorine**, or **pentachloride of antimony** ( $\text{SbCl}_5$ ). Of these, the first is the most important. It may be prepared by dissolving sulphide of antimony in hydrochloric acid, and distilling the liquid. The trichloride passes over, and forms, on cooling, white crystals. It is used for bronzing gun-barrels.

Antimony forms, with sulphur, two compounds corresponding to the oxides, viz., a **sesquisulphide** ( $\text{Sb}_2\text{S}_3$ ) and **antimonic sulphide** ( $\text{Sb}_2\text{S}_5$ ). The former occurs native, as **grey antimony ore**, possessing a bluish black colour, and metallic lustre; and it is the only ore from which the metal is obtained.

Antimony combines with hydrogen, and forms a gaseous compound called **antimoniuretted hydrogen** ( $\text{H}_3\text{Sb}$ ), which is colourless, and burns with a bluish flame, depositing white fumes of antimonious oxide. This gas is produced, when sulphuric acid is added to an antimony salt and zinc.

#### TESTS FOR ANTIMONY.

1. Sulphuretted hydrogen gives an **orange** precip. of antimonious sulphide, which is soluble in sulphide of ammonium.

2. If the above precipitate be washed and gently dried, and a few drops of hydrochloric acid added, trichloride of antimony is formed, which, when thrown into water, gives a white precipitate.

**Blow-pipe test.** Compounds of antimony, mixed with carbonate of soda, give in the inner flame globules of the metal, which may be dissolved in hot hydrochloric acid, and tested as above.

#### (5) ARSENIC.

*Symbol As. Combining weight 75. Sp. gr. 5.7 to 5.9.*

Arsenic is sometimes classed amongst the non-metallic elements. It resembles phosphorus in its chemical properties. It is sometimes found native, but it commonly occurs alloyed with iron, cobalt, nickel, copper, or tin.

The metal is extracted from its ores by roasting them in a furnace, and condensing in flues the sesquioxide of arsenic thus produced. This oxide is further purified and sold as **white arsenic**. The powdered oxide, mixed with charcoal, is heated in a retort,—the metal passes off in vapour, and is condensed in a cool part of the apparatus.

**Properties.** Arsenic possesses a steel-grey colour and brilliant lustre, but it oxidizes in moist air, and becomes tarnished. It is crystalline and very brittle. When heated it passes into vapour, without fusing, which is colourless, and smells strongly of garlic. It burns in air with a bluish flame, forming the sesquioxide. Thrown into chlorine gas, it takes fire, and is changed into the trichloride of arsenic. Nitric Acid changes the metal into arsenic acid, but hydrochloric acid has little action upon it.

Gun shot consists of lead and a little arsenic. The sesquioxide is largely used in preparing green and yellow paints, and also in the manufacture of flint glass.

**OXIDES OF ARSENIC.** There are two oxides,—the **sesquioxide** or **arsenious acid** ( $\text{As}_2\text{O}_3$ ) and **arsenic acid** ( $\text{As}_2\text{O}_5$ ).

**Arsenious acid**—white arsenic ( $\text{As}_2\text{O}_3$ )—is produced by burning the metal in air, and by roasting the ores. It exists in the form of crystals, and also as a glassy substance. It is slightly soluble in water, but dissolves in the alkalis, forming with them soluble compounds called **arsenites**. This oxide forms with copper the pigment called **Scheele's green**, which is a strong poison.

**Arsenic acid** ( $\text{As}_2\text{O}_5$ ) is obtained by the action of nitric acid upon arsenious acid, and evaporating to dryness. It is a white powder, soluble in water; if the solution be evaporated, hydrated crystals of arsenic acid are formed. This acid is decomposed by heat into arsenious acid and oxygen.

It is a powerful acid, and forms with metals salts corresponding to the phosphates, called **arseniates**. The arseniates of the alkalis are soluble in water; those of the other metals are insoluble, but are dissolved by the acids. This acid is more poisonous than arsenious acid.

#### COMPOUNDS OF ARSENIC.

Arsenic and sulphur form three compounds, viz., **realgar** ( $\text{As}_2\text{S}_2$ ); the **sesquisulphide** or **orpiment** ( $\text{As}_2\text{S}_3$ ); and the

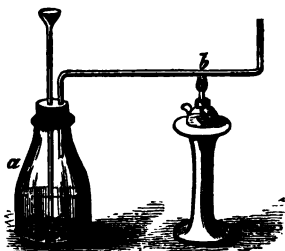
pentasulphide ( $\text{As}_2\text{S}_5$ ). The first two occur native, but they may be prepared artificially by heating arsenic acid with sulphur. The pentasulphide is deposited as a yellow powder, when sulphuretted hydrogen is sent through a solution of arsenic acid.

Arsenic also forms compounds with chlorine, bromine, and iodine. Like antimony, it combines with hydrogen, and forms **arseniuretted hydrogen** ( $\text{AsH}_3$ ), which is a gaseous body of some importance, corresponding to phosphuretted hydrogen. It may be prepared by decomposing an alloy of arsenic and zinc with dilute sulphuric acid. The greatest care is needed in the preparation of this gas, as one bubble has been known to cause death. It burns with a bluish white flame, producing arsenious acid; if a cold body be held in the flame, metallic arsenic is deposited upon it. Chlorine decomposes this gas with flame; it is also decomposed by sending it through a red-hot tube. At a temperature  $40^\circ$  below zero, the gas is reduced to a colourless liquid.

#### TESTS FOR ARSENIC.

1. Sulphuretted hydrogen gives in an acid solution a **bright-yellow** precip. of **orpiment**, soluble in ammonia.
2. Sulphate of copper gives in neutral solutions a **green** precip. of **arsenite** of copper (Scheele's green).

**Marsh's Test.** This test depends upon the production of arseniuretted hydrogen. Zinc and dilute sulphuric acid are put into a flask (a) having a funnel tube, and a delivery tube bent at right angles. The arsenical liquid is added by means of the funnel, and arseniuretted hydrogen is given off. If a lamp be held at (b), the gas, in passing, is decomposed, and deposits metallic arsenic. In



employing this test, it is necessary that the zinc and sulphuric acid be pure.

The remaining metals, on account of their rarity, require but little notice. Some of their compounds are used in the arts and manufactures.

### URANIUM.

*Symbol U. Combining weight 120. Sp. gr. 18.4.*

This metal is sparingly distributed, and is found combined with the two minerals, pitchblende and uranite. It is closely allied in its chemical relations to iron and manganese. It has a steel white colour; does not oxidize in air at ordinary temperatures; burns brilliantly when heated; and dissolves in sulphuric and hydrochloric acids. It forms several oxides and salts. The uranium compounds are useful for glass staining, and photography. The protoxide, or uranous oxide ( $\text{UO}$ ), gives a black colour to glass, whilst the sesquioxide, or uranic oxide ( $\text{U}_2\text{O}_3$ ), colours glass a beautiful yellow.

### INDIUM.

*Symbol In. Combining weight 71.8. Sp. gr. 7.36.*

This metal is found in certain ores of zinc, and was discovered by means of the spectrum analysis. It is a white, malleable metal, slowly soluble in hydrochloric acid; its compounds give to flame a blue colour.

### TITANIUM.

*Symbol Ti. Combining weight 50.*

This is a very rare metal, analogous to tin, and found combined with iron. It forms several compounds, but none of them are used in the arts.

## MOLYBDENUM.

*Symbol Mo. Combining weight 96. Sp. gr. 8.62.*

This is a white, brittle metal, the chief ore of which is the disulphide, found in Bohemia and Sweden. It forms three oxides; one of these, **molybdic trioxide** ( $\text{MoO}_3$ ), acts upon bases and yields salts, called molybdates. The molybdate of ammonia is employed as a test for phosphoric acid acidulated with nitric acid; a yellow liquid is produced, which, on heating, deposits a yellow precipitate. The compounds of molybdenum are not used in the arts.

## TUNGSTEN (Wolframium).

*Symbol W. Combining weight 184. Sp. gr. 17.6.*

This is found combined with lime in the mineral **scheelite**, and with ferrous oxide in **wolfram**. It has been only obtained as a dark-grey powder, which is very hard, and difficult of fusion. A little of the metal mixed with steel gives the latter extraordinary hardness. Tungsten forms several compounds.

## THALLIUM.

*Symbol Tl. Combining weight 204. Sp. gr. 11.9.*

This metal was discovered in 1861, in the deposits of sulphuric acid manufacture. In physical properties it resembles lead. It has a bluish-white lustre; it oxidizes on exposure to the air; it is soft, malleable, and ductile. It melts at  $561^\circ \text{F.}$ ; if heated higher, it burns with a brilliant green flame. It burns in chlorine, and combines directly with bromine, iodine, sulphur, and phosphorus. As it oxidizes in air, it is best preserved in water. Nitric and sulphuric acids dissolve it readily, with evolution of hydrogen. It forms several oxides and salts; the latter are poisonous.

## TELLURIUM.

*Symbol* Te. *Combining weight* 129. *Sp. gr.* 6.65.

This substance, on account of its analogy to sulphur and selenium, is sometimes classed amongst the non-metallic elements; but it is generally found combined with various metals, and it has a metallic lustre. It is chiefly found in Transylvania and Hungary. It resembles bismuth in colour; fuses between 800° and 900° F., and, at a higher temperature, is converted into a yellow vapour. It burns in air with a blue flame edged with green.

It forms with oxygen two oxides; and with hydrogen it forms a gas, (Telluretted hydrogen,  $H_2Te$ ), which cannot be distinguished from sulphuretted hydrogen.

The remaining metals, Niobium or Columbium, Tantalum, Vanadium, Palladium, Rhodium, Ruthenium, Osmium, Iridium, require no special mention.



## GENERAL QUESTIONS ON THE METALS PROPER.

1. Describe the method of extracting Zinc from its ores. How would you detect its presence in a solution. (C.)

2. What are the chemical differences between Wrought Iron, Cast Iron, and Steel? Explain Bessemer's process for the manufacture of Cast Steel. 1.5 grammes of pure iron wire are ignited in excess of Oxygen Gas: what is the nature, and what the weight of the resulting compound? (C.)

3. What are the properties of the metal Platinum? How is it prepared? Name its most important salt, and state its use in the laboratory. (C.)

4. A piece of green paper is supposed to contain Arsenic. How would you try to ascertain whether it does contain it or not? (O.)

5. The atomic weights of Silver and Chlorine being 108 and 35.5, determine the quantity of metal contained in 14 grains of Chloride of Silver. (O.)

6. What reactions will prove the presence of the following Bases in a solution—(1) Lime, (2) Alumina, (3) Cobalt, (4) Sesquioxide of Iron, (5) Peroxide of Tin? (O.)

7. If 50 grains of ignited Oxide of Copper be exposed to a current of Hydrogen until its reduction is completed, how much Copper will be obtained?

The atomic weight of Copper is 31.75. (O.)

8. Describe the metallurgic process by which Lead is obtained from its most common ore. (O.)

9. Ten grains of a mixture of Chloride of Silver and Iodide of Silver contain  $1\frac{1}{4}$  grains of Chlorine. How much Iodine and Silver are present? The atomic weights of Chlorine, Iodine, and Silver are 35.5, 127, and 108. (O.)

10. Explain what is meant by an equivalent of a substance. If 100 parts of Oxide of Copper, reduced by Hydrogen, give 22.69 of water, deduce the equivalent of Copper. (C.)

11. Explain how pure Silver may be made from an alloy of that metal with Copper. (C.)

12. Explain how Sulphate of Copper is prepared, and how Copper may be detected in a mixed solution. (C.)

13. State and explain the effect of a solution of Hydrochloric Acid upon each of the following : (1) Protoxide of Lead, (2) Red Oxide of Iron, (3) Black Oxide of Manganese, (4) Sulphide of Antimony, (5) Silver, (6) Nitric Acid. (C.)

14. How are (1) Cobalt, (2) Arsenic, and (3) Manganese recognised by the Blow-pipe ? (O.)

15. What reactions are characteristic of the presence of the following metals in a solution : (1) Copper, (2) Bismuth, (3) Lead ? (O.)

16. The atomic weight of Iron being 28, determine the quantity of metal contained in 100 grains of Sesquioxide of Iron. (O.)

17. By what means may Metallic Silver be obtained from its Chloride ? (O.)

18. From what ore is Copper most commonly procured ? Describe the process of its reduction. Express by symbols its compounds with Oxygen. (C.)

19. A person at a distance from a town needs some Nitrate of Silver for analytical purposes : he has pure Nitric Acid in his laboratory, how can he improvise the required re-agent ? (C.)

20. What are the compounds of Iron and Oxygen ? What are the chemical characters of the Salts of Iron ? (C.)

21. Define—(1) Metal, (2) an Alloy. What is the composition of (i) Brass, (ii) Pewter, (iii) German Silver ? (C.)

22. In what form does Chromium occur in nature ? How would you convert Chromic Acid into Sesquioxide of Chromium ? What are the chemical characters of these substances ? (C.)

23. Describe Marsh's test for Arsenic, and state the precautions necessary in using it. (C.)

24. 100 grains of a substance are found to contain Iron 63·6, Sulphur 36·4. Find its chemical formula,  $\text{Fe} = 28, \text{S} = 16$ . (C.)

25. Describe any of the methods adopted in the manufacture of Alum. What is its formula ? (C.)

26. What are the principal ores of Iron, and in what districts are they found ? Describe the method by which Cast Iron is converted into Wrought Iron. (C.)

27. What changes do the following substances undergo when heated ? Oxide of Silver, Nitrate of Copper, Nitrate of Ammonium, Sulphate of Iron, Oxide of Calcium, Chloride of Platinum and Potassium. (O.)

28. How are the following metals affected when heated to redness in the air: Iron, Copper, Silver, Bismuth, Tin? (O.)

29. What are the sources, and what is the mode of preparation of the metal Aluminum? (O.)

30. What takes place when Red Lead (Minium) is boiled with Nitric Acid? \*

31. What is the composition of Arsenetted Hydrogen? and how is it prepared? \*

32. Give the name and composition of the mineral in which Mercury occurs in nature, and describe the process by which the metal is obtained? \*

33. The Copper is to be precipitated from 100 grammes of Crystallized Sulphate of Copper ( $\text{CuSO}_4 + 5\text{H}_2\text{O}$ ) by means of Metallic Zinc. How much Zinc is required? \*

34. You are required to convert 1lb. of Silver into Argentic Nitrate (Nitrate of Silver)—how will you do it, and what will be the Weight of Nitrate produced? \*

35. A compound of Iron and Oxygen contains 72.4 per cent. of Iron, and 27.6 per cent. of Oxygen; what is its formula? \*

36. Calculate the percentage composition of each of the following: Cupric Sulphate ( $\text{CuSO}_4 + 5\text{H}_2\text{O}$ ), Alum ( $\text{KA12SO}_4 + 12\text{H}_2\text{O}$ ), Nitric Acid ( $\text{HNO}_3$ ), Potassium Ferrocyanide ( $\text{K}_4\text{FeC}_6\text{N}_6$ ). \*

## HINTS ON QUALITATIVE ANALYSIS.

---

**ANALYSIS** is the name given to that branch of Chemistry which has for its object the discovery of the **composition** of some unknown body. It is subdivided into two branches, viz., **Qualitative** and **Quantitative** Analysis; the former discovers the **nature** and properties of the component parts of a compound, and the latter determines the **quantity** of each present.

The ordinary way of analysing a simple compound,—i.e. a body containing only **one** base and **one** acid, or **one** metal and **one** non-metallic element,—is to mix some **reagent** or **test** with a solution of the substance to be analysed. The **reagent** is so called because, when placed in contact with another substance, it produces a particular appearance or change, termed the **reaction**, by which the presence of a particular substance is shown.

The following are the chief operations in Qualitative Analysis :—

1. **SOLUTION.** This term denotes the perfect union of a solid with a fluid. The liquid which dissolves the body is called the **solvent**. Solutions are of two kinds—**simple** or **mechanical**, and **chemical**. A solution is simple when no chemical change has taken place, as a solution of salt and water; it is chemical when an entirely new substance is formed, consequent upon chemical combination, as chalk dissolved in hydrochloric acid.

When the solvent has received as much of the solid as it can contain, the solution is called **saturated**.

2. **PRECIPITATION.** This is a term given to the formation of an insoluble compound of the body held in solution, suddenly produced by a reagent, or precipitant. The insoluble substance thus formed is called a precipitate. Precipitates are classified according to their appearance, as curdy, crystalline, gelatinous, &c. Precipitates are called cloudy or turbid, when they are so fine and so small as only to affect the transparency of the fluid.
3. **FILTRATION.** This is the separation of the solid matter in a liquid, by passing the latter through suitable filtering paper. The liquid which passes through the filter is called the filtrate.
4. **DECANTATION.** This is the separation of heavy solid particles in a fluid, by simply decanting the vessel and pouring the liquid off.
5. **EVAPORATION.** This is the process of obtaining solid matter from solutions, by evaporating the latter by the aid of heat. If the heat be gently applied, the solid matter frequently takes a crystalline form, and the process is then called crystallization.
6. **DISTILLATION** is the term given to the process of separating a liquid from a solid by the application of heat. The volatile liquid passes in a gaseous form into a receiver, and is there condensed into the liquid state.
7. **IGNITION** does for solid bodies what evaporation does for liquids: it separates the volatile and non-volatile parts of a solid by heating in vessels called crucibles.
8. **SUBLIMATION** is the term given to the distillation of a solid. The volatile part of the solid is driven off by ignition, and is recondensed in a receiver by cold. The substance is then called a sublimate.

9. **FUSION** is the conversion of a solid into the fluid form by heat. When an insoluble solid is fused along with some other substance, in order to form a soluble compound, the process is called **fluxing**, and the body employed for this purpose is termed a **flux**.
10. **DEFLAGRATION** is a term applied to all decompositions attended with a noise or detonation. Nitrates and chlorates deflagrate when heat is applied to them in the "dry state."
11. **THE BLOW-PIPE.** The use of this is a most important operation. A description of the instrument and its uses have been given at page 37.
- 

## LIST OF CHEMICALS FOR ANALYSIS.

- 6 funnels, about  $2\frac{1}{4}$  in. diameter.  
25 cut filters, about 4 in. diameter.  
24 test-tubes, about 6 in. long by  $\frac{5}{8}$  in. wide.  
1 stand for the same.  
2 glass stirring rods, about 9 in. long.  
1 porcelain evaporating dish, about 4 in. diameter.  
1 tripod, or retort-stand, with ring and triangle to support the dish.  
1 spirit-lamp, to contain about 8 oz., provided with wick and spirit.  
1 washing-bottle, to hold about 1 pint.  
1 small piece of platinum-foil, about 2 in. long by 1 in. broad.  
1 piece of platinum wire, about  $2\frac{1}{2}$  in. long.

- 1 blow-pipe.
  - 2 or 3 pieces of charcoal for blow-pipe experiments.
  - 2 or 3 small pieces of hard glass tubing, about 4 in. long and  $\frac{1}{4}$  in. wide, closed at one end.
  - 1 knife or spatula.
  - 1 apparatus for preparing Hydrogen or Sulphuretted Hydrogen.
  - Hydrochloric Acid.
  - Nitric Acid.
  - Sulphuric Acid.
  - Hydro-sulphuric Acid (or apparatus for preparing the same).
  - Tartaric Acid.
  - Ammonia.
  - Chloride of Ammonium.
  - Sulphide of Ammonium.
  - Carbonate of Ammonium.
  - Oxalate of Ammonium.
  - Potash.
  - Ferrocyanide of Potassium.
  - Chloride of Barium.
  - Nitrate of Silver.
  - Phosphate of Sodium.
  - Sulphate of Magnesium.
  - Lime Water.
- 

- Borax.
- Carbonate of Soda.
- Copper.
- Zinc.
- Litmus and Starch Paper.

## DIRECTIONS FOR ANALYSING SIMPLE COMPOUNDS.

---

1. The substance to be examined should be reduced to a fine state of powder. Examination in the dry state by means of the blow-pipe is described in Table (I).
2. In making a solution of the substance, the usual solvents are water and the acids, diluted hydrochloric, diluted nitric, and aqua-regia, in the order named: First, put a little of the powdered body in a test-tube with a little distilled water. If the substance fails to dissolve in the cold, hold the tube over the spirit-lamp, and boil. If it is still undissolved, add **gradually** one of the above acids, beginning with hydrochloric.
3. Examine first for the **base**, and be careful to use the reagents **gradually**, and also consider for what purpose you add the particular reagent. Table (II) shows the classification of the metallic bases, and the tests by which the various groups are distinguished. Tables (III), (IV), (V), (VI), (VII), furnish tests for each base.
4. After finding the base, examine for the **acid** by means of Table (VIII). Some of the acids are found in examining for the base.
5. In using the reagents, if a precipitate is not formed at once, stir with a glass rod, or heat the solution.
6. Use the substance to be analysed with economy. Always keep a portion for confirmatory tests.
7. Be careful to note every change during the process of examination, and immediately make a record of it.
8. The test-tubes, &c., must be perfectly clean.
9. Needless exposure to injurious gases, as chlorine, sulphuretted hydrogen, &c., should be avoided.





#### IV.

Heat the substance in a glass tube, sealed at one end.

1. The substance is unaltered:—Absence of organic bodies, and volatile matter.

2. The substance intusible; { (a.) From white to yellow, } Oxide of Zinc.  
changes colour, and to white again on cooling,

{ (b.) From white to yellowish } Binoxide of Tin.  
brown,

{ (c.) From white to brownish } Oxide of Lead.  
red, turning to yellow on cooling,

{ (d.) From white or pale yellow to orange yellow & reddish brown, } Peroxide of Bismuth.  
turning pale yellow on cooling,

{ (e.) From red to black, } Sesquioxide of Iron.  
turning reddish brown on cooling,

3. The substance volatilizes and sublimates, { (a.) a yellow sublimate, } Sulphur.  
{ (b.) a white " } Ammoniacal Salts.  
{ (c.) liquid globules, } Mercurial Compounds.  
{ (d.) white crystals, } Arsenical " Antimony.  
{ (e.) brilliant needles,

4. Gases escape, { (a.) Oxygen shows the presence of Nitrates, Chlorates, Peroxides,

{ (b.) Sulphurous Acid indicates Sulphates.

{ (c.) Carbonic Acid " Carbonates.

{ (d.) Carbonic Oxide " Oxalates.

{ (e.) Hydrosulphuric Acid " Sulphides.

{ (f.) Ammonia " Ammoniacal Salts.

#### V.

Heat the substance in a test-tube with concentrated Sulphuric Acid.

1. The substance blackens - - - - - Organic matter.

Coloursless, slightly pungent, (may be) Carbonic Acid.  
or reddish brown, Nitric Acid.

" peculiar odour, blackens } Hydrosulphuric  
paper moistened with Ac- } Acid.  
tate of Lead,

" pungent, forming dense } Hydrochloric  
fumes as it leaves the } Acid.  
tube,

" very pungent, forming } Hydrofluoric  
dense fumes, and deposit- } Acid.  
ing a white coating on the  
upper part of the tube,  
(the latter, when cleared,  
is found to be etched),

" strong smell of vinegar, Acetic Acid.

Yellowish green, peculiar odour, and } Chloric Acid.  
possessing bleaching properties, (the  
evolution accompanied by an ex-  
plosion at times),

2. A gas is evolved

**TABLE II.**  
**CLASSIFICATION OF METALLIC BASES.**

Group I. Precipitate produced by Hydrochloric Acid.	Group II. Precipitates produced by Sulphuretted Hydrogen.	Group III. Precipitate produced by Sulphide of Ammonium, after adding Chloride of Ammonium and Ammonia.	Group IV. Precipitate produced by Carbonate of Ammonium in pre- sence of Chloride of Ammonium.	Group V. Bases remaining after the separation of the former groups.
<p>—</p> <p>METALS.</p> <p>LEAD—Chloride.</p> <p>SILVER—Chloride.</p> <p>MERCURY—{ Subchloride.</p> <p>The above precipitates are white.</p>	<p>Division I.</p> <p>Sulphides insoluble in Sulphide of Am- monium.</p> <p>—</p> <p>MERCURY—black i p.</p> <p>LEAD—black pp.</p> <p>BISMUTH—black i p.</p> <p>COPPER—black pp.</p> <p>ADMIUM—{ ye w pp.</p> <p>PLATINUM—{ brow- in black pp.</p> <p>GOLD—{ brown: sh black i p.</p>	<p>Division II.</p> <p>Sulphides soluble in Sulphide of Am- monium.</p> <p>—</p> <p>TIN—{ brown pp. The bisul- phide yellow.</p> <p>ANTIMONY—{ orange red.</p> <p>ARSENIC—yellow pp.</p> <p>IRON—black pp.</p> <p>ALUMINUM—{ white pp.</p> <p>CHROMIUM—{ green pp.</p> <p>ZINC—white pp.</p> <p>MANGANESE—{ flesh col- oured.</p> <p>COBALT—black pp.</p> <p>NICKEL—black pp.</p> <p>N.B.—The three first metals are pre- cipitated also by Am- monia and Chloride of Ammonium.</p>	<p>—</p> <p>BARIUM.</p> <p>STRONTIUM.</p> <p>CALCIUM.</p> <p>The above are pre- cipitated as Carbon- ates, and are of a white colour.</p>	<p>—</p> <p>MAGNESIUM.</p> <p>POTASSIUM.</p> <p>SODIUM.</p> <p>AMMONIUM.</p> <p>—</p> <p>N.B.—Phosphate of Ammonium separ- ates Magnesium.</p>

TABLE III.

## BASES GROUP I.

To the solution add HCl in excess. If there is a precipitate, it is either silver chloride ( $\text{AgCl}$ ), mercurous chloride ( $\text{HgCl}$ ), or lead chloride ( $\text{PbCl}_2$ ).

To distinguish between these, take a small part of the wet precipitate, and pour on it an excess of ammonia ( $\text{NH}_3$ ).

Redissolves, $\text{AgCl}$ .	Turns black, $\text{HgCl}$ .	Is unaltered, $\text{PbCl}_2$ .
Confirm.	Confirm.	Confirm.
Add Nitric Acid ( $\text{HNO}_3$ ); the precipitate reappears.	A drop of the original solution placed on a bright slip of copper ( $\text{Cu}$ ), will, in a few minutes, produce a spot which, when gently rubbed, appears shining like silver.  If the copper is gently heated, the mercury ( $\text{Hg}$ ) volatilizes, and the silvering disappears.	The dried precipitate heated on charcoal before the blow-pipe, gives malleable metallic globules.  Sulphuric Acid ( $\text{H}_2\text{SO}_4$ ) produces a white precipitate of sulphate of lead ( $\text{PbSO}_4$ ).

TABLE IV.

## BASES GROUP II.

To the solution treated with HCl add Hydrogen Sulphide (Sulphuretted Hydrogen,  $H_2S$ ), until it smells strongly after shaking. If only a white precipitate is formed, proceed as if there were none. If there is a coloured precipitate, it may be one of the following: PbS, HgS, Bi<sub>2</sub>S<sub>3</sub>, CuS, CdS, SnS, SnS<sub>2</sub>, Sb<sub>2</sub>S<sub>3</sub>, As<sub>2</sub>S<sub>3</sub>, Au<sub>2</sub>S<sub>3</sub>, Pt<sub>2</sub>S<sub>3</sub>.

*To determine which of these is present, observe the colour.*

Yellow.		Orange.		Brown or Black.	
<p>May be CdS, As<sub>2</sub>S<sub>3</sub>, or SnS<sub>2</sub>.</p> <p>To determine which mix a portion of the fluid (with the precipitate) with excess of <math>NH_3</math>, add 4 or 5 drops of <math>(NH_4)_2HS</math>, and warm.</p>		<p>Sb<sub>2</sub>S<sub>3</sub>.</p> <p>—</p> <p>Confirm.</p>		<p>PbS, HgS, Bi<sub>2</sub>S<sub>3</sub>, CuS, SnS.</p> <p>Test the original solution (a small portion at a time) with the following, (except in the last case, when the dark precipitate is used):—</p>	
<p>Undissolved.</p> <p>CdS.</p> <p>—</p> <p>Confirm.</p>	<p><i>Disolved.</i></p> <p>As<sub>2</sub>S<sub>3</sub>, SnS<sub>2</sub>.</p> <p>To the original solution add <math>NH_3</math>.</p>	<p>H<sub>2</sub>SO<sub>4</sub>.</p> <p>A white precipitate.</p>	<p>KHO.</p> <p>A yellow precipitate.</p>	<p><math>NH_3</math>, in excess.</p> <p>If on adding <math>NH_3</math> the precipitate at first formed is redissolved, leaving a blue solution.</p> <p>Cu.</p> <p>Confirm.</p> <p>Test original solution with K ferrocyanide—a brownish red precipitate.</p>	<p><math>NH_3</math>, in excess, and then 5 or 6 drops of <math>(NH_4)_2HS</math>, and warm.</p> <p>—</p> <p>This being done to the fluid containing the dark precipitate, it will dissolve if it is</p>
	<p>White precipitate.</p> <p>Sn.</p> <p>Confirm.</p>	<p>Pb.</p> <p>Hg.</p>	<p><math>NH_3</math>, in excess.</p> <p>A white precipitate.</p> <p>Filter; dissolve the precipitate in a drop or two of HCl; and then add plenty of water. The solution becomes milky.</p> <p>Bi.</p>	<p>Confirm.</p> <p>Add to this solution some diluted HCl, a light yellow precip.</p>	<p>Sn</p> <p>Confirm.</p> <p>Add to this solution some diluted HCl, a light yellow precip.</p>

TABLE V.  
BASES GROUP III.

To the original solution add Chloride of Ammonium and Ammonia, and, whether the latter reagent produces a precipitate or not, a little Sulphide of Ammonium, and apply heat. A precipitate shows the presence of one of the following: Fe, Co, Ni, Mn, Zn, Cr, Al, or Silicic Acid.

*Observe the Colour.*

Black.		Flesh-coloured.		Bluish Green.		White.	
<i>May be the Sulphides of Fe, Co, or Ni.</i>		<b>Mn.</b>		<b>Cr.</b>		<i>May be Al, Zn, or Silicic Acid.</i>	
To distinguish these, add to the original solution Potass or Soda.		Confirm with Blow-pipe.		Confirm with Blow-pipe.		To distinguish these, add to the original solution a few drops of Soda.	
Dirty greenish white precip., changing to reddish brown on exposure to air.	Light green precip.	Confirm with Blow-pipe.		Confirm with Blow-pipe.		A precipitate which dissolves in excess indicates— Zn or Al,	No precipitate. Silicic Acid.
	<b>Ni.</b>						
<b>Fe.</b>	Confirm with Ammonia.					Add Hydrogen Sulphide.	
	Confirm with Ferrocyanide of Potassium.						
						A white precipitate.	No precipitate.
						Zn.	Al.
						Confirm with Blow-pipe.	Confirm with Blow-pipe.



**TABLE VIII.**  
**TO DETECT THE PRESENCE OF THE MOST IMPORTANT ACIDS.**

I.				
Hydrochloric Acid is added to the solution.				
(a.) If effervescence ensues, Carbonic Acid or Hy. Sulfuric Acid is present. These may be easily distinguished by the smell of rotten eggs, which the latter acid possesses.				
(b.) If, on adding Chloride of Barium to the Acid Solution, a white precipitate is formed, insoluble in acids, it indicates				
<b>Sulphuric Acid.</b>				
(c.) If, on heating the Acid Solution, it changes to a green colour, and evolves Chlorine Gas, the presence of Chromic Acid is indicated.				
(d.) Evaporate the Acid Solution to dryness, and mix the residue with Hydrochloric Acid and water; if an insoluble residue remains, it indicates				
<b>Silicic Acid.</b>				
II.				
Acidify the solution with Nitric Acid and add Nitrate of Silver; if a white precipitate be formed (Chloride of Silver), which dissolves in Ammonia, there is present				
<b>Hydrochloric Acid.</b>				
(b.) If Chlorine water imparts to the original solution a yellow tint, it indicates the presence of				
<b>Bromine.</b>				
(c.) If the precipitate be not dissolved by Ammonia, and possesses a pale yellow colour, it indicates the presence of				
<b>Iodide.</b>				
III.				
Add to the original solution Chloride of Ammonium, Ammonia, and Sulphate of Magnesium. Agitate the mixture, and if a white crystalline precipitate be formed, it indicates the presence of				
<b>Phosphoric Acid.</b>				
The precipitate in this case is—				
<b>Ammonia-phosphate of Magnesium.</b>				
IV.				
The substance is heated in a test-tube with concentrated Sulphuric Acid; the formation of dense white fumes, depositing a white film on the damp side of the glass, which etches the glass, indicates the presence of				
<b>Hydrofluoric Acid.</b>				
(b.) If copper filings be added to the mixture in the tube, and ruddy fumes of Nitrous Oxide be evolved on the application of heat, there is present				
<b>Nitric Acid.</b>				
(c.) If Alcohol be added to the substance mixed with Sulphuric Acid, then ignited, and stirred while burning, a green flame indicates				
<b>Boric Acid.</b>				
(d.) If on heating, the pungent odour of Acetic Acid is given off, it indicates				
<b>Acetic Acid.</b>				
V.				
The substance heated with Hydrochloric Acid evolves a yellow gas, having a sweet smell, indicates the presence of				
<b>Chloric Acid.</b>				
<i>N.B.</i> —Chlorates and Nitrates heated on Charcoal, deflagrate violently.				



## APPENDIX TO TABLES OF ANALYSIS.

As Table II necessitates the almost constant use of Hydrogen Sulphide, the following course shows how several metallic bases may be discovered without the use of this reagent.

To a solution of the substance, add a few drops of **Sulphide of Ammonium**.

<b>No precipitate</b> indicates an Alkali, or an Alkaline Earth.		<b>A precipitate.</b> <i>Observe the colour.</i>	
<b>No precipitate</b> indicates one of the Alkalis.		<b>White</b> indicates <b>Zinc or Aluminium.</b>  To distinguish, add to the original solution, Caustic Potash in excess, and then Chloride of Am- monium.	<b>Flesh colour</b> indicates <b>Manganese.</b>
<b>No precipitate</b> indicates one of the Alkalis.		<b>A white precipitate</b> indicates one of the Alkaline Earths.	<b>Black</b> indicates one of the remaining <b>Metals proper.</b>
To distinguish, <i>Vide</i> Table VII.		<b>A white precipitate</b> indicates Aluminium.	To distinguish, proceed as in Table II.
To distinguish, <i>Vide</i> Table VI.		<b>No precipitate</b> indicates Zinc.	

## COMPARISON OF THE METRICAL WITH THE COMMON MEASURE.

## MEASURES OF LENGTH.

	In English Inches.	In English Feet = 12 Inches.	In English Yards = 3 Feet.	In English Fathoms = 6 Feet.	In English Miles = 1,760 Yards.
Millimetre . . . . .	0.03937	0.0032809	0.0010936	0.0005468	0.0000006
Centimetre . . . . .	0.39371	0.0328090	0.0109363	0.0054682	0.0000062
Decimetre . . . . .	3.93708	0.3280899	0.1093633	0.0546816	0.0000621
Metre . . . . .	39.37079	3.2808992	1.0936331	0.5468165	0.0006214
Decametre . . . . .	393.70790	32.8089920	10.9363310	5.4681655	0.0062138
Hectometre . . . . .	3937.07900	328.0899200	109.3633100	54.6816550	0.0621382
Kilometre . . . . .	39370.79000	3280.8992000	1093.6331000	546.8165500	0.6213824
Myriometre . . . . .	393707.90000	32808.9920000	10936.3310000	5468.1655000	6.2138244

1 Inch = 2.539854 Centimetres.  
1 Foot = 3.0479449 Decimetres.

1 Yard = 0.91438348 Metre.  
1 Mile = 1.6093449 Kilometre.

## MEASURES OF SURFACE.

	In English Square Feet.	In English Sq. Yards = 9 Square Feet.	In English Poles = 272.25 Sq. Feet.	In English Roods = 10.800 Sq. Feet.	In English Acres = 43.560 Sq. Feet.
Centiare or sq. metre . . . . .	10.7642993	1.1960338	0.0395383	0.000988457	0.0002471143
Are or 100 sq. metres . . . . .	1076.4299342	119.6033260	3.9538290	0.098845724	0.0247114310
Hectare or 10,000 sq. metres . . . . .	107642.9934183	11960.3326020	395.3829399	9.884572398	2.4711430996

1 Square Inch = 6.4519689 Square Centimetres.  
1 Square Foot = 9.2903043 Square Decimetres.

1 Square Yard = 0.83609715 Square Metre or Centiare.  
1 Mile = 0.404671021 Hectare.

## MEASURES OF CAPACITY.

	In Cubic Inches.	In Cubic Feet = 1,728 Cubic Inches.	In Pints = 34.6593 Cubic Inches.	In Gallons = 8 Pints = 277.27384 Cubic Inches.	In Bushels = 8 Gallons = 2218.19075 Cubic Inches.
Millilitre, or cubic centimetre . . .	0.061027	0.0000353	0.001761	0.00022010	0.000027512
Centilitre, or 10 cubic centimetres . .	0.610271	0.0003532	0.017608	0.00220097	0.000275121
Decilitre, or 100 cubic centimetres . .	6.102705	0.0035317	0.176077	0.02200967	0.002751208
Litre, or cubic decimetre . . . . .	61.027052	0.0353166	1.760773	0.22009668	0.027512085
Decalitre, or centistere . . . . .	610.270515	0.3531658	17.607734	2.20096677	0.275120846
Hectolitre, or decistere . . . . .	6102.705152	3.5316581	176.077341	22.00966767	2.751208452
Kilolitre, or stere, or cubic metre . .	61027.051519	35.3165807	1760.775314	220.09667675	27.512084594
Myriolitre, or decastere . . . . .	610270.515194	353.1658074	17607.734140	2200.96676750	275.120845937
1 Cubic Inch = 16 3861759 Cubic Centimetres.					
1 Gallon = 4.54547969 Litres.					
Cubic Foot = 23.318119 Cubic Decimetres.					

## MEASURES OF WEIGHT.

	In English Grains.	In Troy Ounces = 480 Grains.	In Avordupois Lbs. = 7,000 Grains.	In Cwt. = 112 Lbs. = 784,000 Grains.	Tons = 20 Cwt. = 16,820,000 Grains.
Milligramme . . . . .	0.015432	0.000032	0.0000022	0.00000002	0.000000001
Centigramme . . . . .	0.154323	0.000323	0.0000220	0.00000020	0.000000010
Decigramme . . . . .	1.543235	0.003215	0.0002205	0.00000197	0.000000098
Gramme . . . . .	15.432349	0.032151	0.0022046	0.0001968	0.000000984
Decagramme . . . . .	154.323488	0.321507	0.0220462	0.0019684	0.000009842
Hectogramme . . . . .	1543.234880	3.215073	0.2204621	0.0196841	0.000098421
Kilogramme . . . . .	15432.348800	32.150727	2.2046213	0.01968412	0.000984206
Myriogramme . . . . .	154323.488000	321.507267	22.0462126	0.19684118	0.009842059
1 Grain = 0.06479896 Gramme.					
1 Troy oz. = 31.103466 Gramme.					
1 lb. Avo. = 0.45359235 Kilogr.					
1 Cwt. = 50.80237089 Kilogr.					

# INDEX.

## A

	PAGE
Absolute weight.....	2
Acid oxides.....	8
how named.....	9
Agate.....	56
After-damp, or choke-damp.....	30
Air, composition of, by weight and volume.....	20
Air, a mechanical mixture.....	20
physical properties of.....	22
Alabaster.....	62
Alkalis, metals of the.....	67-68
distinguishing characters of the.....	78
Alkaline earths, metals of.....	67-78
in combination, tests for.....	84
Alloys.....	66
Alloys of zinc.....	92
lead.....	109
silver.....	113
tin.....	128
Alum, its composition and manufacture.....	69
Alumina, its nature and properties.....	88-89
Aluminum.....	88
compounds.....	89-90
compounds, tests for.....	90-91
bronze, Abyssinian gold.....	88
Amalgam.....	66
Amalgamation, method of.....	112
Amethyst.....	56
Ammonia, composition of.....	75
preparation and properties of.....	75-76
Ammonia in the air, and its use to plants.....	21
Ammonia produced in manufacture of coal gas.....	34
Ammonium and its compounds.....	75-77
tests of.....	77
Amorphous bodies.....	65
phosphorus.....	59
Analysis, qualitative and quantitative.....	135
Analysis, operations in.....	135-137

	PAGE
Analysis, necessary apparatus, &c. 137-138	
Analysing simple compounds, directions for.....	139
Analysis, tables of.....	140-148
Anhydrous bodies, definition of.....	16
Animal charcoal.....	23
Antimony, properties of.....	124
compounds of.....	125
tests for.....	126
Aqueous vapour in air.....	21
Aqua regia.....	41
Aquafortis.....	22
Archimedes and specific gravity.....	2
Arithmetical exercises, rules for.....	86
Arsenic, properties of.....	126
compounds of.....	127
tests for.....	128
Arsenuretted hydrogen.....	128
Atmosphere, composition of, &c.....	20
Atomic weight.....	3
theory.....	3
volume of gases in relation to hydrogen.....	17
Azote.....	19

## B

Barium, properties of.....	78
salts of.....	79
Baryta, protoxide of barium.....	78
Bell-metal.....	128
Basic oxides.....	8
Bessemer's steel process.....	95
Bismuth, properties of.....	117
compounds.....	117
tests for.....	118
Bittern, or mother liquor.....	44
Black oxide of manganese.....	99
Black-jack.....	91
Black lead, or plumbago.....	27
Blast furnace, construction of.....	94
Bleaching properties of chlorine.....	30
Bleaching powder.....	82
Blende.....	91
Blowpipe, description of.....	37



# INDEX.

153

	PAGE
Corrosive sublimate .....	116
Crith the, and uses of .....	17
Crystallization .....	65
water of .....	16
Crystal systems .....	65
Cubic nitre .....	73
Cupric oxide .....	107
Cupreous oxide .....	106

## D

Davy Lamp .....	36
Decantation .....	136
Decomposition of bodies .....	1
in gaseous	
form, shown .....	86
Decomposition of water by elec-	
tricity .....	15
Deflagration .....	137
Density .....	2
of gases .....	17
of compound gases .....	18
Diamond, nature and properties of	
Didymium .....	27
Didymium .....	91
Dimorphous bodies .....	65
Distillation .....	136
Distilled water .....	17
Distribution of metals .....	67
Ductility of metals .....	65
Dutch liquid .....	34

## E

Earth's crust, composition of .....	56
Earthenware, manufacture of .....	90
Earths proper, metals of .....	67
Ebullition of water .....	16
Element, definition of .....	1
Elementary bodies, list of .....	4
Emery .....	88
Epsom salts, or sulphate of mag-	
nesium .....	83
Equivalent, chemical .....	3
Erbium .....	91
Evaporation .....	136
Expansion of gases by heat .....	11
Exceptions to the law of volumes ..	18

## F

Fahrenheit's scale .....	10
Fast colours given by alumina .....	89
Felspar .....	83
Ferric acid .....	96
Ferric salts .....	97
Ferrous salts .....	96
Filtration, process of .....	136
Fire damp in mines .....	32
Flame, nature of .....	36

	PAGE
Flowers of sulphur .....	49
Fluoride of silicon .....	43
Fluorine, and its compounds .....	47
Fluor-spar .....	47
Flux, and Fluxing .....	137
Formula of a body given to find	
per-centage composition .....	85
French weights and measures .....	149
Freezing point of water .....	16
Furnace for iron smelting .....	94
Fusion, process of .....	137
Fusible metal .....	117
Fusibility of metals .....	64

## G

Galena, chief ore of lead .....	108
Galvanized zinc .....	92
Gas coal, manufacture of .....	34
Gas-burner, Bunsen's .....	36
Gaseous volumes, measurement and	
weight of .....	17
Gases, expansion of, by heat .....	11
relation of volume to pres-	
sure .....	11
Gases, combination and decompo-	
sition of .....	86
General questions on alkalis and	
alkaline earths .....	86
General questions on metals proper	
General properties of metals .....	132
German silver .....	64
German silver .....	92
Glass, chief constituent of .....	57
etching upon, with hydro-	
fluoric acid .....	47
Glauber's salt .....	72
Glucinum .....	91
Gold, occurrence of .....	119
properties of .....	119
compounds of .....	120
tests for .....	120
Jeweller's .....	119
British standard .....	119
dust .....	119
Graphite, or plumbago .....	27
Green vitriol .....	97
Gun-metal .....	123
Gunpowder, composition of .....	70
Gypsum .....	82

## H

Hæmatite, or specular iron ore ....	93
Hardness of water, cause of .....	81
Hartshorn, spirits of .....	75
Heavy spar .....	79
Horn silver .....	113
Hydrate of potash .....	69
definition of .....	16
Hydraulic mortar .....	81



N		PAGE			PAGE
Neutral oxides .....	8		Phosphites .....	61	
Nickel .....	102		Phosphoric acid .....	61	
— compounds of .....	103		Phosphorous acid .....	60	
— tests for .....	103		Phosphorus .....	58	
Niobium .....	131		Phosphorus, properties of .....	59	
Nitre, or saltpetre .....	70		— oxygen compounds .....	60	
Nitric acid .....	22		Phosphuretted hydrogen .....	61	
— anhydrous .....	23		Photography, use of silver salts in .....	113	
— test for .....	147		— use of sodium salt in .....	113	
Nitrous oxide .....	24		Plaster of Paris .....	82	
Nitric oxide .....	25		Platinum .....	120	
Nitrous acid .....	25		— properties of .....	121	
Nitric peroxide .....	26		— compounds of .....	121	
Nitrogen .....	19		— tests for .....	122	
— preparation of .....	19		Plumbago .....	27	
— properties of .....	20		Porcelain .....	90	
— in air, uses of .....	21		Pot-metal .....	109	
— questions on .....	26		Potashes, or pearl-ashes .....	70	
Nordhausen sulphuric acid .....	51		Potash, caustic .....	69	
O			Potassium .....	68	
Oil of vitriol .....	51		— preparation of .....	69	
Olefant gas .....	33		— salts of .....	70	
Onyx .....	56		— tests for .....	71	
Opal .....	56		— carbonate of .....	70	
Ores of metals .....	67		— chlorate of .....	71	
Orpiment .....	127		— nitrate of .....	70	
Organic chemistry, definition of .....	1		Precipitant, definition of .....	136	
Osmium .....	131		Precipitate, definition of .....	136	
Oxides, how obtained .....	8		Precipitation, process of .....	136	
— classification of .....	8		Priestley, discovery of oxygen .....	6	
— how named .....	8		Protoxide, definition of .....	8	
Oxidating flame of blow-pipe .....	37		Proportion, combining .....	8	
Oxygen, discovery of .....	6		Puddling, process of .....	94	
— preparation of .....	6		Purple of Cassius .....	120	
— properties of .....	7		Pyrites, iron .....	96	
— called vital air of atmo- sphere .....	7		Q		
Oxygenated water .....	17		Qualitative analysis .....	135	
Oxyhydrogen flame .....	16		Quantitative analysis .....	135	
Ozone in air .....	9		Quartz .....	56	
— properties of .....	10		Quicklime .....	80	
P			Queen's metal .....	123	
Palladium .....	131		Questions upon oxygen, &c. ....	12	
Paris, plaster of .....	82		— hydrogen .....	18	
Pearl-ashes .....	70		— nitrogen .....	26	
Perchloric acid .....	42		— carbon, &c. ....	32	
Peroxide, definition of .....	8		— compounds of car- bon and hydrogen .....	38	
Periodic acid .....	47		Questions upon flame and blowpipe .....	38	
Peroxide of hydrogen .....	17		— chlorine and com- pounds .....	43	
Permanganic acid .....	100		Questions upon iodine, &c. ....	48	
Pewter .....	123		— sulphur and its compounds .....	54	
Phosphates .....	61		Questions upon phosphorus and its compounds .....	63	
			Questions upon alkalis and alkaline earths .....	86	
			Questions on metals proper .....	132	



## R

	PAGE
Rain-water .....	16
Radicals, simple salt .....	39
Realgar .....	127
Reaumur's scale .....	10
Reagent, definition of .....	135
Red-fire .....	80
Red-lead .....	110
Red phosphorus .....	59
Reducing flame of blowpipe .....	37
Relation of gases to pressure .....	11
Rhodium .....	131
River-water .....	16
Rock crystal .....	56
Rubidium .....	74
Ruby .....	88
Ruthenium .....	131

## S

Safety lamp .....	86
Salt-ammoniac .....	75
Salt, definition of .....	9
— how named .....	9
— common .....	72
Salt-radicals .....	89
Salt-petre, or nitre .....	70
Sand in mortar, its use .....	81
— white, composition of .....	55
Sapphire, composition of .....	88
Scheele's green .....	127
Sea-water, composition of .....	16
Selenic acid .....	55
Selenious acid .....	55
Selenium .....	54
Selenite .....	82
Seleniuretted hydrogen .....	55
Sesqui, definition of .....	9
Silica .....	56
Silicic acid .....	56
Silicon .....	56
Silver .....	111
— extraction of .....	112
— pure, how obtained .....	112
— properties of .....	112
— alloys of .....	113
— compounds of .....	113
— tests for .....	114
Small shot .....	109
Smelling-salts .....	77
Soda-ash .....	73
Soda caustic .....	72
Sodium .....	71
— preparation of .....	72
— compounds of .....	72
— tests for .....	74
Solution .....	135
Speculum metal .....	123
Specific gravity, definition of .....	2
— how to find .....	2

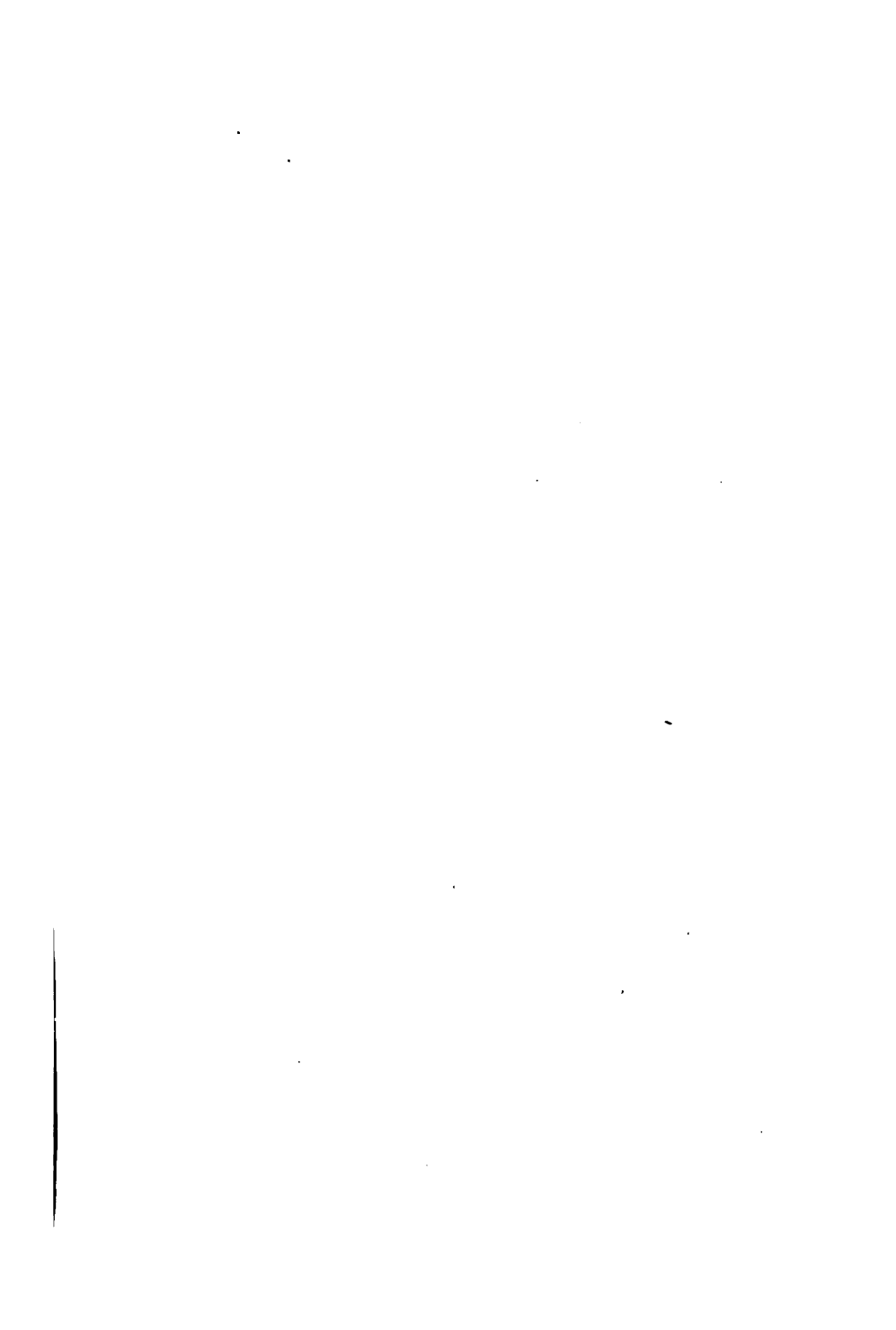
	PAGE
Specific gravity, how discovered by	
Archimedes .....	2
Spirits of hartshorn .....	75
Stalactites .....	81
Standard gold .....	119
— silver .....	113
Stannic oxide .....	123
Stannous oxide .....	123
Stannous salts .....	123
Steam, weight of one litre .....	18
Steel, manufacture of .....	95
— tempering of .....	95
Strontia .....	80
Strontianite .....	80
Strontium .....	79
— salts of .....	80
Structure of flame .....	35
Stucco .....	82
Sublimation .....	136
Sugar of lead .....	110
Sulphides, metallic .....	63
Sulphur .....	48
— how obtained .....	48
— properties of .....	49
— and oxygen compounds .....	49
Sulphuric acid .....	50
— Nordhausen .....	51
— tests for .....	52
Sulphurous acid .....	49
Sulphur and hydrogen compounds .....	52
— and carbon, compound of .....	54
— and compounds, questions .....	54
on .....	54
Sulphuretted hydrogen .....	52
— tests for .....	53
Sunlight, effect upon plants .....	21
Symbols of elements .....	3-4
Synthesis, example of process of .....	29
Systems of crystallography .....	65

## T

Table of blow-pipe analysis .....	140
Tables of analysis .....	142-148
Table of elements .....	4
Tables of weights and measures .....	149
Tantalum .....	131
Tar .....	84
Tartar emetic .....	125
Tellurium .....	131
Temperature, measurement of .....	10
Tempering of steel .....	95
Terbium .....	91
Tetrachloride of phosphorus .....	43
Thallium .....	130
Thermometers .....	10
Thorium .....	4
Tin .....	122
— preparation of .....	122
— properties of .....	122
— alloys of .....	123

	PAGE		PAGE
Tin, compounds of .....	123	Water of crystallization .....	16
— tests for .....	124	Water, hardness of, explained ....	16, 82
Tin-stone .....	122	Weight, absolute .....	1
Tin-plate .....	123	— atomic .....	2
Titanium .....	129	— combining .....	3
Tungsten, or Wolframium .....	130	— relative .....	3
Type-metal .....	109	— specific .....	2
<b>U</b>		Weights, tables of .....	149
Uranium .....	120	White arsenic .....	126
<b>V</b>		White lead .....	110
Vanadium .....	131	Witherite .....	79
Vapour density, how to find .....	17	Wolframium .....	130
Vermilion, ingredient of .....	115	Wood-charcoal .....	28
Vitriol, oil of .....	50	<b>Y</b>	
— blue .....	107	Yttrium .....	91
— green .....	97	<b>Z</b>	
— white .....	92	Zero-point on thermometer .....	10
Volatile alkali .....	75	Zinc .....	91
<b>W</b>		— blende .....	91
Water .....	15	— extraction of .....	91
— properties of .....	16	— properties of .....	92
— varieties of .....	16	— salts of .....	92
		— tests for .....	93
		— alloys of .....	92
		Zirconium .....	4









1





